PHYSICAL SCIENCE

FOR

MATRICLUATION CLASSES

PART II

CHEMISTRY

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PREFACE

THIS little book forms Part II of *Physical Science* written primarily for the young students preparing for the Matriculation and School Leaving Certificate Examinations of the University of the Panjab. Chemistry is essentially an experimental science; an endeavour, therefore, has been made throughout the book to include the largest number of simple experiments (one hundred and thirty-three) which can easily be performed by the teacher in the classroom or by the student himself with very little of special apparatus. The subject matter has been presented in a simple and systematic manner. Indian names of certain substances and indigenous methods of their preparation have been given wherever possible. At the end of each chapter has been added a summary of the main points discussed in it so as to help to fix on the memory the chief facts of the lesson.

It is hoped that the book will be found useful by those for whom it is written. Suggestions and criticisms for the improvement of the book from science teachers will be received with thanks.

AUTHORS

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PREFACE TO THE FOURTH EDITION

It is a matter of great pleasure that the Science Teachers have found the book to be of some use. The book has been recommended by the Panjab and Delhi Universities as a Text-book for the M. & S.L.C students. Efforts have been made to make the language simpler. Certain other improvements have been made. No work being perfect, there is room for further improvement. We, therefore, request the Science Teachers using the book to please inform us of the drawbacks and short comings that come to their notice. We shall by their co-operation try to make the work as nearly perfect as possible.

AUTHORS

PREFACE TO THE SEVENTH EDITION.

The book has been thoroughly revised brought uptodate. In certain cases hints for cessful performance of experiments have been Nearly all the experiments given in the book been successfully performed in the laboratory. chapter on chemical formulae and equations h. added. Some scientific hobbies which can be taken up by students have also been added. hope the teachers will find the present work useful. Any further suggestions to improve the are right welcome.

AUTHORS

CHAPTER I.

PHYSICAL AND CHEMICAL CHANGES.

It is a matter of common observation that bodies often undergo temporary changes under certain conditions. Thus water on cooling freezes into hard, brittle ice; but the ice melts and becomes water on warming. A bar of iron which has been heated to redness is slightly changed, and becomes the same on cooling. Again, an iron knitting needle when rubbed with a bar-magnet acquires the property of attracting iron filings towards it but loses this property when hammered.

In other cases the changes may be more deep-seated. Thus, milk may become sour, iron when exposed to damp air rusts into reddish-brown powder, a candle burns away and apparently disappears. When the white of an egg is gently heated it changes from a transparent, colourless liquid to an opaque white solid.

In the first case the changes are only slight and temporary and are said to be Physical changes; while changes in the second case are much more drastic, resulting in the disappearance of the original substance and the formation of a new substance. They are called Chemical [illegible]es.

The following experiments further illustrate the important facts about these changes.

Physical changes

EXPERIMENT 1. Heat a little powdered sulphur in a clay crucible. It melts into a yellow liquid. On heating further, the liquid becomes thick dark in colour. Continue heating, the thick liquid again becomes thin. On cooling, the above changes are repeated in the reverse order.

EXPERIMENT 2. Heat a platinum or iron wire in a flame It becomes red hot and begins to glow. When removed from the flame it regains its original appearance and remains unchanged.

EXPRIMENT 3. Weigh a beaker. Put some water in it, and add two grams of powdered common salt in the water. Stir the contents of the beaker, the salt disappears in the water. Heat the contents of the beaker till the water is driven out and a dry powder is left behind. Cool and then weigh it. The weight of the salt remains unchanged.

In all these cases the composition of the substances remains unchanged and no new substance is produced. The change is easily reversed by altering the outside conditions; the weight has remained unchanged. *Thus a physical change is a temporary change in which the composition of the body undergoing the change, remains unaltered*

Chemical changes.

EXPERIMENT 4. Hold a piece of paper or cloth in a flame. It burns away leaving behind a black ash.

EXPERIMENT 5. Place a piece of magnesium ribbon in a flame. It burns with a dazzling light and leaves behind a white residue.

EXPERIMENT 6. Put some copper turnings in a test tube and pour some nitric acid over them. A reddish brown gas will be produced and no copper turnings will be left in the test tube after some time.

EXPERIMENT 7. Pass electric current through acidulated water contained in a voltameter. The water is broken up into oxygen and hydrogen.

EXPERIMENT 8. Heat a piece of lead in a clay crucible. It melts into a shining liquid. Continue heating strongly; stirring all the while. The metallic lead changes into a yellow powder.

EXPERIMENT 9. Heat a few crystals of lead nitrate in a test tube. Reddish-brown fumes are given off and a residue is left behind in the tube.

In all these cases new substances with new properties are produced ; the composition of the substance in each case is changed. The change is not easily reversed by merely altering the outside conditions. The weight of a substance changes when it is changed into a new material. *Thus a chemical change is more or less a permanent change in which a new substance with different properties is formed*

The Law of Conservation of Matter. Although a Physical or a Chemical change can bring about a temporary or permanent change in the composition and properties of matter, it can neither create nor destroy matter. The quantity of matter in the universe remains constant, it does not increase or decrease. In a chemical change if the whole of the products are collected, they are found to be equal in weight to the substances entering into the reaction. The following experiments c[illegible]ly illustrate the above facts.

EXPERIMENT 10. In a dry clean flask put a small quantity of dry, cold sand. Introduce a piece of dry phosphorus with a pair of tongs and fit an air-tight cork in it. Weigh quickly the whole apparatus. Now gently warm the flask holding it over a small flame. Stop heating when the phosphorus starts burning. The flask is soon filled with dense white fumes. Allow the flask to cool and weigh again. There is no change in the total weight: the phosphorus has disappeared

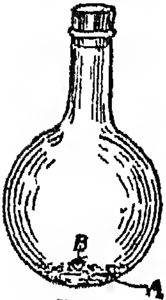


Fig. 1.
A–sand
B–phosphorus.

EXPERIMENT. 11. Take a little potassium iodide solution in a test tube and place it inside a conical flask containing solution of lead acetate. The flask is corked and weighed. Tilt the flask so as to mix the liquids. Notice a beautiful yellow coloured substance is formed. Re-weigh the flask There is no change in the weight although a chemical change has taken place.

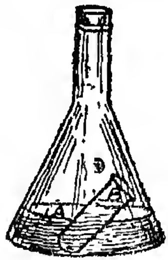


Fig 2
A–solution of lead acetate.
B–solution of potassium iodide

It is clear that in these two experiments there has been no creation or destruction of matter, but merely a redistribution. In fact innumerable experiments conducted by chemists have shown that when any chemical or physical change takes place the total weights of the substances formed are exactly

qual to the sum of the weights of the original substances. These facts are summed up in the Law of Conservation of Matter which states that during chemical or physical changes matter is neither destroyed nor created. The principle is also known as the Law of Indestructibility of Matter.

SUMMARY

A temporary change in which the composition of the body undergoing the change remains unaltered is called a physical change

A more or less permanent change in which a new substance with different properties is formed, is called a chemical change

The law of Conservation of matter. During chemical changes matter is neither destroyed nor created

QUESTIONS.

1 What do you understand by a physical change ? Describe some simple experiments to illustrate your answer

2 What is chemical change ? Illustrate by examples.

3 Distinguish clearly between a Physical and a Chemical change Give examples to illustrate your answer

4 State, with reasons, whether the changes taking place during the following cases are physical or Chemical —

(1) Rusting of iron (2) Freezing of ice cream (3) Magnetizing iron (4) Burning a candle (5) Burning an electric lamp (6) Electrolysis (7) Baking bread (8) Dissolving sugar in milk

5 What Chemical and Physical changes are observed to take place when the following substances are strongly heated ? —

(1) A crystal of salt (2) A crystal of alum (3) A bright piece of copper. (4) Magnesium ribbon (5) Piece of sodium (6) Piece of phosphorus (7) Piece of chalk (limestone) (8) Powdered sulphur. (6) Ammonioum chloride

6. State giving reasons, whether the changes involved in the following are Physical or Chemical —

(1) Heating of red oxide of mercury (2) Freezing of water (3) Adding water to quick lime (4) Heating the white of an egg (5) Converting water into steam (6) Heating lead nitrate (7) Burning of sulphur (8) Change of milk into curd

7 Describe two epxeriments to illustrate the Law of Conservation of matter State the law

CHAPTER II.

LABORATORY OPERATIONS.

In the laboratory we make use of certain simple mechanical processes for the purification and separation of substances. The most important of them are:

Solution. It is a matter of common experience that when a solid such as sugar is put into water it disappears and seems to become a liquid. Sugar is, however, not lost but has spread throughout the water and its presence may be proved by the sweet taste that it has given to the water. In the same way, if a little salt is added to some water, it also disappears and gives to the water its own taste. When sugar or salt so completely spreads throughout water, it is said to have dissolved in water. The process of one substance dissolving in the other or the product obtained is known as a *solution* The substance which dissolves is called the *solute* and the substance in which it is dissolved is called the *solvent*. Thus, in the above cases, water is the solvent and sugar or salt is the solute.

A solution contains the dissolved substance or solute in an extremely finely divided condition so much so that it is not possible to detect the presence of the dissolved particles with the naked eye or even by the help of a microscope. The articels of the dissolved substance are uniformly distributed through the liquid and do not settle down on long standing

EXPERIMENT. 12 Into a small beaker containing distilled water put some powdered chalk (limestone) and shake the liquid round. The water becomes milky in colour, because the particles of chalk float about in it. Allow the beaker with its contents to stand for some time. The chalk particles settle down at its bottom leaving a clear liquid above. Take a little of this clear liquid on a watch glass and heat it to dryness. Notice no residue is left behind.

The above experiment shows that chalk does not dissolve in water. Such substances as do not dissolve in water are called insoluble substances. Substances like sugar alum, common salt which dissolve in water are called soluble substances. Substances that dissolve readily in water are sugar, copper sulphate, salt, alum, borax, dyes, etc. Among the insoluble substances may be mentioned sand, powdered glass, wax, chalk, sulphur, rubber, wood, paper, phosphorus, charcoal, resin, oil, sealing wax, etc.

Other solvents. Water is not the only solvent, all liquids possess the power of dissolving solids to a greater or lesser degree. Some substances that do not dissolve in water are readily soluble in other liquids.

EXPERIMENT 13. Shake up a little quantity of powdered sulphur with some water in a test tube. Notice that sulphur does not dissolve in water. Into another test tube containing powedered sulphur add little carbon disulphide drop by drop and shake up. Sulphur is seen to have dissolved completely in carbon disulphide.

EXPERIMENT 14. Add a little powdered sealing wax to methylated spirit contained in a dish and stir

up with glass rod. Sealing wax soon dissolves in the spirit

Oil and grease are readily soluble in benzene or petrol which is used in dry-cleaning soiled clothes. Resin, sealing wax, camphor and iodine dissolve in methylated spirit. Sulphur and yellow phosphorus dissolve in carbon disulphide.

Solution of liquids in liquids Water dissolves not only solids but some liquids as well. In a test tube containing water add a little alcohol and shake it up. The two liquids dissolve in each other completely. On the other hand oils do not dissolve in water. Mercury, turpentine, carbon disulphide, chloroform are other examples of liquids that do not mix with water; while honey, milk, phenyle, glycerine, etc. dissolve in water. Either dissolves partially in water.

Solution of gases in liquids. Gases are also soluble in water, but some dissolve more than others. The gases such as, oxygen, nitrogen, air, are slightly soluble in water; carbon dioxide, sulphur dioxide are soluble; while ammonia gas, hydrochloric acid gas are extremely soluble in water. Hydrogen, carbon monoxide and coal gas are practically insoluble in water.

Saturated and unsaturated solutions

EXPERIMENT 15. Half fill a small beaker with water and add to it a little of powdered nitre and stir with a glass rod. Nitre dissolves in the water. Add a little more and stir the solution; nitre again disappears. Continue adding nitre till no more will dissolve and some remains at the bottom.

Now heat the solution in the beaker over a flame. The nitre at the bottom of the beaker dissolves. Add more nitre until some remains undissolved at the bottom. Heat the solution further; the undissolved nitre goes into solution again. Continuing in this way and by heating it the solution can dissolve more nitre. Allow the solution to cool. Some of the nitre separates out of solution and settles at the bottom of the beaker.

From the above experiment we learn that if we continue to add nitre to the cold water in the beaker; a point is reached at which some of it remains undissolved in that quantity of water. The water in the beaker can no longer dissolve any more solid. When no more of a substance can dissolve in a liquid at a particular temperature; the solution is said to be saturated at that temperature. If, on the other hand, the solvent can dissolve more of the soluble substance the solution is called unsaturated.

When the saturated solution is heated, it becomes unsaturated, *i.e.*, more of the solid can dissolve in it. Generally the hotter the solvent the more solid will it dissolve. On allowing a hot saturated solution to cool, some of the soluble solid separates out and settles at the bottom.

Solubility. We have seen already that some substances like sugar, nitre, dissolve readily in water; others, such as lime, dissolve only slightly; while glass, sand, etc., are insoluble. The amount of a solid that can be dissolved in a liquid depends not only on the nature of the substance, but also on the temperature. Solids, generally, dissolve more in hot water than in cold water. Lime and sodium sulphate anhydrous, however, dissolve more in cold water than in hot water.

The solubility of some salts, *e g*, nitre, increases rapidly as the temperature rises. Again common salt is equally soluble in hot as well as in cold water. Its solubility is 35·9.

The amount of a substance that can be dissolved in 100 *parts of water at a particular temperature is called the solubility of that substance at that temperature.* It can be determined as follows :—

EXPERIMENT 16 Take some water in a beaker and add a little common salt to it and shake well. If the salt dissolves add more and shake again. Repeat the addition, if necessary, till some remains undissolved at the bottom. Filter (*See* Fig. 6) the saturated solution thus obtained. In a clear dry dish which has been previously weighed take some of the clear solution. Weigh the dish with the solution. Note the temperature of the solution. Now, heat gently the dish over a flame till whole of this water is driven out. Cool the dish and weigh it again, and thus get the weight of the dry salt left in the dish

Calculate the solubility of salt as below :—

Temperature of solution	$=25^{\circ}$C.
Wt. of evaporating dish	$=48·50$ gm.
Wt. of evaporatiog dish +solution	$=69·00$ gm.
∴ Wt. of the solution	$=20·50$ gm.
Wt. of the dish and of the residue	$=54\ 00$ gm.
Wt. of the dish	$=48·50$ gm.
∵ Wt. of the residue	$=5·50$ gm.
Wt. of water	$=15·00$ gm.

Hence 15·00 gm. of water can dissolve 5·5 gm. of common salt at 25°C.

$$\therefore \text{Solubility of salt at } 25^{\circ}\text{C.} = \frac{5{\cdot}5}{15{\cdot}00} \times 100 \text{ gm.}$$
$$= 36{\cdot}6.$$

EXPERIMENT 17. Fit up the apparatus as shown in Fig. 3. Fill the flask and the delivery tube completely with water. Heat the flask, bubbles of a gas

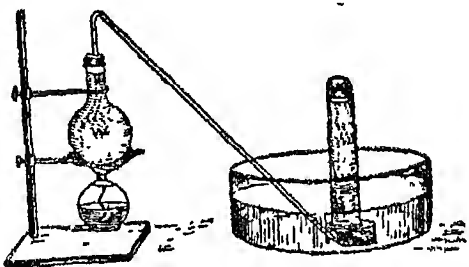


Fig 3 W—water. A—air.

come out and are collected in the test tube. This gas is found to be common air.

EXPERIMENT 18. Open a soda-water bottle. Notice the rise of bubbles of a gas which escape at the mouth. These are the bubbles of carbon dioxide which has been dissolved in water under pressure, but which now escapes when the pressure is decreased.

The above experiments show that :—

1. Gases dissolve less in hot water than in cold water ; and
2. Gases dissolve more under increased pressure.

Evaporation. If water is heated or exposed to air in an open vessel, it gradually disappears in the form of vapour. This process by which a liquid disappears on exposure to air or on heating is known as *Evaporation.*

EXPERIMENT 19. Weigh out a quantity of salt in an evaporating basin and dissolve it in water. Heat it gently over a burner. To ensure gradual heating, a sand bath may be used (Fig 4). A sand bath generally consists of an iron tray containing some sand in it. The water boils and evaporates away completely. A white solid remains in the basin. Weigh again. It is found that the weight is equal to the weight of the basin and salt before solution and that the solid left is still salt.

Fig 4

A—evaporating basin.

B—sand bath.

Soluble substances like sugar; salt, etc., may be obtained from their solutions by evaporation.

Common salt for domestic purposes is obtained from the sea water by evaporation by persons living on the sea coast.

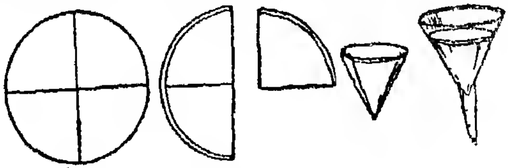
Decantation Insoluble substances like chalk and sand, if shaken with water and allowed to stand, settle down at the bottom and the clear liquid above can be poured off or siphoned off into another vessel. This process is called Decantation.

In most villages of the Punjab where well water is scarce canal water or pond water is filled in pitchers and allowed to stand when the mud settles at the bottom and clear water decanted off and used for drinking purposes. Decantation is also made use of by *niyaryas* in separating particles of gold and silver from the dust and sweepings from a goldsmith's shop.

Filtration. In order to remove completely the insoluble solids or suspended impurities from liquids filtration is usually employed. The process of filtration consists in allowing the liquid containing the insoluble impurities to pass through a porous material like filter paper, coarse cloth, etc., when the clear liquid passes through and leaves the insoluble impurities behind. The clear liquid so obtained is called the *filtrate* and the insoluble impurities left behind form the *residue*. The following experiment

generally carried out in the laboratory illustrates the process of filtration.

EXPERIMENT 20. Take a circular piece of filter paper and fold it in two in the form of a semi-circle, Fig. 5 (ii)



(i) (ii) (iii) (iv) (v)

Fig 5.

then again in the form of a quadrant. Fig. 5 (iii). Open the folded paper to form a hollow cone with three folds on one side and one on the other. Fig. 5 (iv). Place a glass funnel on a filter stand and put this cone in the funnel. Sprinkle a little clear water to make it stick to the walls of the funnel. Place a clean beaker below the funnel. Pour gently the muddy water in the funnel by allowing it to run slowly down a glass rod, the lower end of which just touches the filter paper on the thicker side as shown in Fig. 6. The clear liquid along with dissolved substance, if any, passes down through the filter paper into the beaker and insoluble impurities are left on the filter paper as a residue.

Fig 6
R—glass rod.
M—muddy water.
F—filtrate.

Filter papers are used in laboratories and hospitals where small volumes of liquids are required to be filtered. When filtration on a large scale is to be carried on other substances are used as filters. Coarse cloth, powdered charcoal and sand are commonly employed in daily life. In places where pure water cannot be obtained for drinking purposes, the muddy water is made fit for drinking, in the following simple way :

Three earthen vessels (*gharras*) are placed, one above the other, the upper two having holes at the bottom (Fig 7). The uppermost vessel contains muddy water which gradually passes through sand contained in the middle one. Most of the suspended impurities are held back by the sand, and clear water collects into the third vessel.

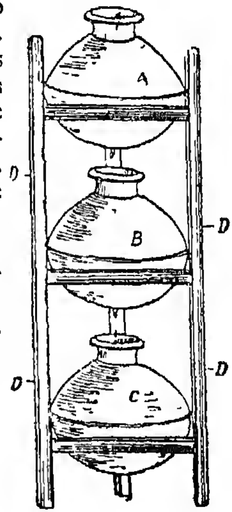


Fig. 7

A–muddy water.
B–vessel containing sand, gravel etc.
C–pure water.
D–frame.

Sand filters are used also for removing the insoluble and suspended impurities from water in big cities like Lahore and Lyallpur. Water under pressure is made to pass through filter beds made of sand and gravel

Several types of ready-made filters are sold in the market. The commonest of these are the Berkfield filter and the Pasteur–Chamberland filter. Each consists of inner unglazed porcelain cylinder through which water is allowed to pass.

Distillation Decantation or filtration can be used only when the impurities present in a liquid are either insoluble or suspended. To remove soluble impurities from a liquid the method known as distillation is used. Distillation consists in the process of changing a liquid into vapour, and collecting and condensing the vapour in another vessel In the laboratory, distillation is carried on as illustrated in the following experiment.

EXPERIMENT 21. Arrange the apparatus as shown in Fig. 8. The apparatus consists of a flask for boiling the solution, a Leibig's condenser, and a flask to receive the distilled liquid.

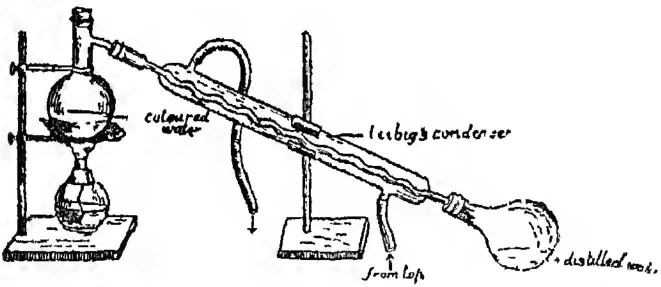


Fig 8.

In the flask put some copper sulphate solution and heat it over a flame. Allow a steady current of cold water to pass through the condenser. As the heating proceeds drops of clear liquid are seen collecting in the receiver. Do not boil to dryness or the flask may crack. The clear liquid collected in the receiver will be free from any kind of impurity. This can be verified by evaporating a drop on a watch glass when no residue will be left behind. This liquid is known as the *distillate*.

If Leibig's condenser is not to be used, a retort may be employed; and the receiver is kept cool by dipping it in cold water. (*See* Fig. 9.)

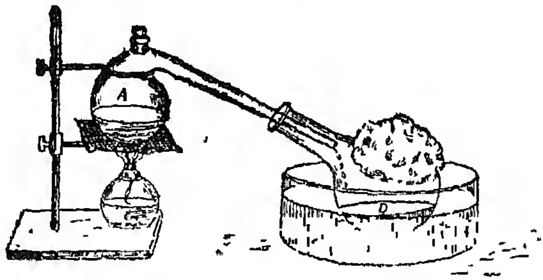


Fig 9

A—retort. D—receiver kept in cold water.

The process of distillation is used in purifying

Fig. 10.

liquids from dissolved impurities. The druggists use

this process for obtaining extracts of flowers and plants. An Indian druggist employs the apparatus shown in Fig. 10 for the preparation of extracts (*araq*).

Fractional Distillation. A mixture of liquids may be separated into its constituents by this process. Thus if a mixture of alcohol and water be distilled it is found that the boiling point in the beginning is 84° C. The liquid that collects in the receiver is richer in alcohol. If the distillation is stopped when one-fourth of the mixture has distilled over, and the distillate is again distilled as before, we get a distillate which is still richer in alcohol. This partical separation of liquids is known as Fractional Distillation.

Distillation is used to obtain pure water for drinking purposes from sea water and other impure waters not fit for drinking. Distilled water used in laboratories and hospitals is also obtained by the process of distillation.

Crystallization We have already learnt that solids dissolve more in hot than in cold water; that is, the solvent power of water increases with rise of temperature. A hot saturated solution contains more of the solute than a cold one. When the hot saturated solution is allowed to cool some of the dissolved substance separates out in the form of regular geometrical figures, known as crystals. This process is known as Crystallization. Crystals have shining faces and sharp edges.

EXPERIMENT 22. Prepare a hot saturated solution of alum: filter it and allow it to cool slowly. After a short time some of the alum is seen to have separated out in the form of crystals.

Crystals may also be obtained by allowing the solvent to evaporate slowly. The slower the

evaporation, or the more gradual the cooling, the larger and more perfect are the crystals obtained.

In nature crystals are also formed by the slow and undisturbed solidification of a melted solid. Crystals found in igneous rocks are formed by this process.

Crystallization is used to purify substances such as nitre or alum or sugar; etc. A strong solution of the substance is prepared in hot water and filtered. The solution is allowed to cool. The impurities present remain in solution while the substances separate out in pure crystals.

Two soluble substances such as nitre and salt, which have different degrees of solubility, held in solution, can be separated by crystallization as the crystals of the one much soluble in hot water will separate out first on cooling. This particular process is called the Fractional Crystallization. In the present case, nitre being more soluble in hot than in cool water, crystallizes out first on cooling.

Sublimation. Certain solids such as camphor, sal-ammoniac, iodine when heated pass off directly into vapour, without changing to the liquid condition and the vapour when cooled, directly changes into solid. This process is known as sublimation. The substance is said to sublime or to form a sublimate.

EXPERIMENT 23. In a large dry flask put a crystal of iodine and fit it with a cork. Gently warm the flask on a low flame. The iodine at once sublimes into a beautiful violet vapour filling the flask. At the cooler parts of the flask the violet vapour condense again into solid iodine.

EXPERIMENT 24. Heat a little powdered ammonium chloride in a dish. Invert over it a cold glass funnel as in Fig. II. Plug the funnel by cotton wool.

The ammonium chloride soon passes off into white vapour which deposits on the cold funnel.

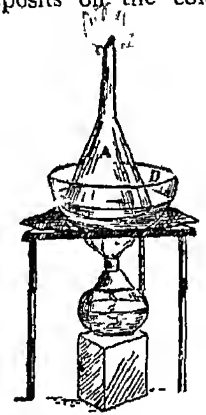


Fig 11.

D—glass funnel. A—dish.

Sublimation is used for separating substances which sublime from impurities present in them.

SUMMARY.

The process of one substance dissolving in the other and the product obtained is known as Solution The substance which is dissolved is known as Solute and the liquid in which it is dissolved is called Solvent

Other solvents besides water are spirit, carbon disulphide benzene (for paraffin wax), etc.

As a rule solids dissolve more in hot than in cold water The solubility of gases decreases with rise of temperature

When a solution has got the maximum amount of solute dissolved in it at a given temperature, the solution is said to be a Saturated Solution at that temperature,

The amount of substance which can be dissolved in 100 parts of a liquid at a given temperature is called its solubility in that liquid at that temperature.

Decantation is the pouring of the clear liquid at the surface when a liquid containing insoluble impurities is allowed to stand and become clear.

Filtration is process of separating insoluble impurities from a liquid by passing it through a porous material

Distillation is the conversion of a liquid into vapour and condensation of vapour into the liquid again

Fractional distillation is the process used to separate two liquids from their mixture

When a hot saturated solution is allowed to evaporate gradually or cool slowly, the excess of the substance is deposited in the form of crystals The process is known as crystallization

Fractional crystallization is the process of separating two soluble substances from a liquid

Sublimation is the direct passage of a solid into vaporous state and *vice versa*

QUESTIONS

I Explain the terms—solute, solvent and solution Give examples,

2 What is a saturated solution ? How would you prepare a saturated solution of nitre at the room temperature ?

3. Define the term solubility How would you determine the solubility of alum at the ordinary temperature ?

4 Upon what does the solubility of a substance depend ? What is the effect of cooling a hot saturated solution ?

5 How will you prepare large and good crystals of copper sulphate ? Give the precautions you will observe ? What are the practical uses of this process ?

6 Describe experiments to show that solids dissolve more in hot water than in cold water while gases dissolve less in hot water than in cold water ? Give examples of solids which behave otherwise

7 Explain the terms—decantation, sublimation, sublimate, distillate, filtrate, residue, evaporation Give examples.

8 Describe the process of distillation as carried out in the laboratory To what practical uses are filtration and distillation put ?

9 A white powder is given to you How will you proceed to find out whether it is soluble in water or not ?

10 Distinguish between decantation and filtration, evaporation and sublimation How will you obtain pure water from a muddy specimen containing salt dissolved in it ?

CHAPTER III.

ELEMENTS, COMPOUNDS AND MIXTURES.

Elements and Compounds. Pure substances are divided into elements and compounds. It is necessary to understand the difference between two classes of pure material substances. The red solid formed by heating mercury in air is a pure substance, but when heated further, it splits up into mercury and oxygen; it is not a simple, but a compound substance. But mercury itself may be heated, it may be boiled indeed or treated in any other way, but it cannot be split up into two other substances. In fact, no means are known by which mercury can be changed into a totally different substance except by combining something else with it. This combination is, of course, a process of addition, and, therefore, the fresh substance formed will weigh more than the original mercury. Any different substance formed from mercury is less simple than mercury. A number of substances are known which resemble mercury in this one point; by no known means can they be separated into two or more substances simpler than themselves. They cannot be made to split up as the substance formed by the combination of mercury with oxygen does when heated. They can only be made into different substances by causing something else to combine with them, *e.g*, magnesium can be made to burn, in other words, to combine with oxygen, and thus from a white powder. The white powder differs from magnesium in that it contains oxygen as well.

Substances like mercury and magnesium, which cannot be separated into simpler substances by any known method, are called *Elements* Substances formed by combining elemens are called *Compounds*. There are about 85 elements known to the modern chemists. The most important of them are :—

Table of common elements.

Solids .

Aluminium.	Gold.	Zinc.
Antimony.	Iodine.	Platinum.
Arsenic.	Iron.	Potassium.
Barium.	Lead.	Radium.
Bismuth.	Magnesium.	Silicon.
Calcium.	Manganese.	Silver.
Carbon.	Nickel	Sulphur.
Cobalt.	Phosphorus.	Tin.
Copper.		Tungsten.

Liquids :

Bromine.
Mercury.

Gases :

Argon.	Hydrogen.
Chlorine.	Nitrogen.
Helium.	Oxygen.

Compounds and mixtures Compounds must be carefully distinguished from mixtures. Every drop of water has exactly the same properties and so is the case with every particle of red oxide of mercury. In case of water, oxygen and hydrogen of which it is composed, and in case of red

oxide of mercury, oxygen and mercury, cannot be recognized in them. On the other hand, if we mix flour and salt and grind them together, we get a mixture in which the substances exist quite distinct from each other. We cannot see the particles separately with naked eyes but by a lens or a microscope the particles of different substances can be easily made out. The following experiments further illustrate the distinction between compounds and mixtures.

EXPERIMENT 25. Mix some iron filings with sulphur and grind them thoroughly in a mortar to obtain a grey coloured mass. Take it out on a paper and examine with a lens. The particles of sulphur and iron filings will be seen lying side by side. Now insert a magnet in the mixture. Iron filings will stick to it and can be removed by moving the magnet in the mixture. Sulphur is thus left behind.

Take a little of the mixture in a dish and add carbon disulphide to it. Sulphur will dissolve in the liquid ; filter it. Iron filings will be left on the filter paper as residue. On evaporation of the filtrate, sulphur will be left behind.

EXPERIMENT 26. Prepare some more mixture of the same substances and put it in a crucible which is supported on a clay pipe triangle on the tripod stand (Fig. 12). Heat the crucible. Observe that the sulphur melts and some burns away. Add more sulphur and again heat the mixture until no more sulphur burns away ; repeat this three or four times. A dark coloured mass will be obtained. Remove the black mass thus obtained and examine it with a lens. The particles of sulphur and iron cannot be recognized. Insert a magnet in the powdered mass. No iron filings can be removed. Put a little of the

substance in a dish containing carbon disulphide and shake well. It does not dissolve in it and

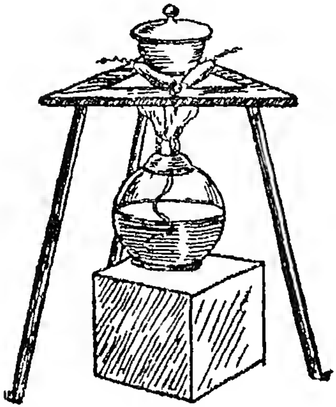


Fig. 12.

sulphur cannot be removed. There is, therefore, neither free iron nor free sulphur.

From the above experiments the following points of distinction between a chemical compound and a mechanical mixture are clear :—

In a mixture the constituents are found to be lying side by side and can be separated by simple mechanical means. The constituents retain their individual properties and the properties of the mixture will be found to be midway between those of its constituents.

In a compound, the constituents cannot be seen lyidg side by side, nor can they be separated by simple mechanical means. Many experiments have shown that when compounds are formed their constituents combine in fixed proportions only and if an excess

of a substance is used, it remains unchanged. The properties of a compound are entirely different from those of its constituents.

Mixture	*Compound*
1. The constituents of a mechanical mixture may be present in any ratio.	The constituents of a chemical compound are always present in a certain definite proportion.
2. The properties of a mechanical mixture are intermediate between those of its constituents.	The properties of a compound are entirely different from those of its constituents.
3. The composition and properties of a mixture may vary in different parts and the constituents may be seen lying side by side.	The composition and properties of a compound are exactly the same throughout. The constituents cannot be seen lying side by side even with the help of a microscope.
4. The components of a mixture can be easily separated by simple mechanical means (such as solution, filtration, decantation, distillation, etc.)	The constituents of a compound cannot be easily separated. They can be separated only as a result of a chemical change.
5 A mixture is formed without any energy change.	When a chemical compound is formed energy in the form of heat, light, or electricity, etc., is either used up or produced.

Separation of a mixture into its constituents The methods commonly used in separating mixtures into their constituents are best shown by the following examples.

Gun-powder. It is a mixture of potassium nitrate (75 parts, powdered charcoal (15 parts), and powdered sulphur (10 parts). Of the three constituents of gun-powder, nitre is very soluble in hot water, sulphur and charcoal are insoluble. But of these two sulphur dissolves in carbon disulphide. Hence to separate them proceed as follows:

EXPERIMENT 27. Shake the gun-powder with sufficent hot water, and filter. The filtrate contains the nitre, which can be recovered by the evaporation of the water. Now wash the residue on the filter paper by pouring water over it a few times, then dry it carefully; remove it from the paper and add carbon disulphide. Shake and allow it to stand. Filter the liquid from the black charcoal and allow it to evaporate slowly. The sulphur is thus recovered. The charcoal remains behind as residue on the filter paper and can be obtained by drying.

Separation of nitre from salt and insoluble impurities Shake well some of the mixture containing nitre, salt and sand in hot water contained in a beaker, filter it. The filtrate is a clear liquid containing nitre and salt in solution. The insoluble substance such as sand is left as residue on the filter paper. Heat the filtrate in a dish to saturate the solution. Now allow it to cool. The nitre crystallizes out first as it is much more soluble in hot than in cold water whereas, salt is equally soluble both in hot and cold water. The liquid left is rich in salt. (Fractional crystallization).

Types of chemical action When two or more substances combine to form a new compound or when a compound breaks up into two or more simpler substances, a chemical action is said to have taken place. *Chemical action is divided into four types* .—

1. *Combination.* Two or more elements combine to form a definite chemical compound, *e.g.*,

Iron+Sulphur=Iron sulphide.

2. *Decomposition.* A compound breaks up or decomposes into two or more elements or simpler compounds, *e g.*,

Mercury oxide=Mercury+oxygen.

Calcium carbonate=Calcium oxide+carbon dioxide,
(chalk) (lime)

3. *Chemical replacement.* One element replaces another in a compound, *e.g.*,

Iron+copper sulphate=Copper+iron sulphate

EXPERIMENT 8. Put bright iron nails in a strong solution of copper sulphate contained in a beaker. After a short time the nails are covered with a reddish layer of copper, and iron goes into solution.

4. *Double decomposition.* Two compounds act upon each other chemically by exchange of parts to produce two new compounds, *e.g.*,

Sodium carbonate+calcium chloride=
(Washing soda)

Sodium chloride+calcium carbonate.
(chalk)

EXPERIMENT 29. Prepare solutions of sodium carbonate and of calcium chloride in two separate

test tubes. Pour one solution into the other ; a white insoluble substance (chalk) is formed and settles at the bottom. Sodium chloride remains in solution as may be found by filtration and evaporation.

Precipitation In the above case (Expt. 29) two soluble substances have combined to form two new compounds, one of which is insoluble and settles at the bottom. This process is known as ***Precipitation*** and the white insoluble substance formed is called the *Precipitate*.

Chemical affinity. Certain substances combine with one another very readily at ordinary temperature and form new compounds, while some substances do not combine at all or do so under special conditions. Phosphorus when exposed to air combines with oxygen of the air and begins to burn while silver or gold if left exposed to air, remain unaffected. Phosphorus may be said to have great attraction for oxygen of the air while silver or gold has none. This attraction which makes the chemical union of one element with another to take place easily, is known as *chemical affinity*. Thus phosphorus has chemical affinity for oxygen while silver or gold has no chemical affinity for it.

Chemists believe that in compounds the elements are held together by this force of chemical affinity.

Conditions and agencies that bring about chemical actions.

EXPERIMENT 30. In a dry test tube put a little powder of mercuric chloride and add to it a small quantity of powdered potassium iodide. Mix the powders ; nothing happens. Now add water, a beautiful

red-coloured precipitate of mercuric iodide is formed :

Mercuric chloride+Potassium iodide=
Mercuric iodide+Potassium chloride

EXPERIMENT 31. In two separate test tubes prepare solutions of lead acetate and potassium iodide. Mix the two. Notice the formation of a yellow precipitate.

EXPERIMENT. 32. Mix some iron filings with powdered sulphur in a mortar. No action takes place, only a greyish mixture is obtained. Heat the mixture in a test tube. A black compound iron sulphide is formed.

EXPERIMENT 33. Pass a current of electricity in copper sulphate solution contained in a voltameter. Copper separates out and deposits on the cathode and the solution becomes acidic.

EXPERIMENT 34. Powder separately a little of antimony sulphide and potassium chlorate. Mix the two gently on a piece of paper. Place the mixture on a stone and suddenly strike it with a hammer. It explodes accompanied by a chemical change

EXPERIMENT 35. Fill a hard glass bottle half with chlorine and half with hydrogen, and put it exposed to the sunlight. The gases combine chemically with a great explosive noise

In each of the above experiments a chemical action has taken place with the help of some external agencies which are :

1. *Close contact.* This is the most essential condition of a chemical action ; and is brought about

by mixing the substances in the form of solutions (Expt. 28, 29).

2. Heat. (Expt. 30).

3. Electricity. (Expt. 31).

4. Pressure. (Expt. 32).

5. Light. (Expt. 33).

SUMMARY

Pure substances are divided into elements and compounds

An element is a substance that cannot be separated into simpler substances by any known method,

A compound is a substance formed by the combination of two or more elements In a compound the components cannot be separated by simple methods The properties of a compound are generally quite different from those of its constituents which are always present in certain definite proportions which for each compound are fixed

A mixture is made of constituents which lie side by side and can be separated by easy mechanical methods The constituents may be present in any proportion and the properties of a mixture are midway between those of its components

The classes of chemical action

1 Chemical combination 2 Chemical decomposition
3 Chemical replacement 4. Chemical double decomposition

Chemical actions are brought about by

1 Close contact.
2 Heat
3 Light,
4 Electricity.
5 Pressure

Chemical affinity The attraction which facilitates the chemical union of one element with another is known as the chemical affinity

Separation of nitre and salt :—

Substance which are less soluble in water than others separate out first from a solution on crystallization,

The solubility of nitre is more in hot water than that of salt but it is less in cool water than that of salt The solubility of nitre at 650 C is 120 but at 150 C it is 25 only whereas salt shows no variation, So when a hot saturated solution of a mixture of nitre and salt is cooled nitre being less soluble in cold water than common salt separates out first

QUESTIONS

1 Explain the terms—element, compound and mixture. Give examples Say to which class each of following belongs —

Chalk, sulphur, diamond, ice-cream, gunpowder, sand, glass, iron filings, water, air, charcoal, soap, brass, ink, milk.

2 Distinguish clearly between a mixture and a compound. Illustrate your answer with experiments

3 Name the types of a chemical action Support your answer with experiments

4 Describe the conditions and agencies which help to bring about a chemical action Give examples in each case

5 You are given a mixture of powdered nitre, sulphur, and charcoal How will you separate the constituents ? Give experimental details

6 Your are given a mixture of common salt, nitre and sand How will you separate them ? State clearly the principle of this method.

7 How will you proceed experimentally to separate the constituents of a mixture of ammonium chloride, nitre and powdered glass ? Give precautions

8 Given a mixture of iron filings, sulphur, copper sulphate, alum and chalk, how will you separate them ?

9 You are given an impure sample of nitre containing small amounts of soluble and insoluble impurities Describe with full practical details, the method that you would adopt in order to obtain its pure and dry crystals

10 Explain as clearly as you can the meanings of the term 'chemical affinity'

11 How will you separate the constituents of gun powder ?

CHAPTER IV.

AIR.

The atmosphere, commonly known as air is a gaseous envelope that surrounds the earth. It is supposed to extend to a height of about 100 miles. The presence of air can be felt in several ways. When we move a fan we feel the current of air striking our body. A vessel which we call empty is in reality full of air. That it is so can be shown by inverting an empty tumbler in a trough full of water ; the water does not rise up. If we tilt it a little, bubbles of air are seen coming out and water rises up inside the vessel.

Air, in fact, is a form of matter, occupies space, and has weight. It has been calculated that due to its weight, air exerts pressure equal to 35 pounds on one square inch of surface.

The Composition of Air. Air is a mixture of several gases, the most important being oxygen and nitrogen. Other gases which are generally present in air are carbon dioxide, water vapour, argon, ozone, oxides of nitrogen and ammonia. The composition of air differs slightly in different places. The air in towns where there are several factories and workshops is much impurer than the air of the countryside. The average percentage composition by volume of air is given below :—

Oxygen ...	... 20·90%	by volume.
Nitrogen ...	... 78·00%	,,
Argon (Neon, Krypton)	... ·94%	,,

Carbon dioxide	...	·03% by volume.
Water vapour	...	*variable.*
Ozone, ammonia, oxides of nitrogen, nitric acid	...	*traces.*

Besides these there are present in the air dust particles, and living germs of several kinds of diseases. Oxygen, nitrogen with other mixed gases are more or less constant constituents of air, whereas the amount of water vapour, carbon dioxide, etc., varies at different places, and at different times.

That nitrogen and oxygen are found to be present in the air in the ratio of 79 to 21 by volume (and 77 to 23 by weight) can be shown by a chemical action, such as the burning of phosphorus, or the rusting of iron, in a known volume of air.

Burning of Phosphorus.

EXPERIMENT 36. Place a piece of yellow phosphorus on a crucible lid resting on a piece of cork which floats on water in a large trough. Invert an uncorked bell jar over it; and mark the level of the water in it by a strip of gummed paper. Divide the space above the water level into five equal parts. (*See* Fig. 13.)

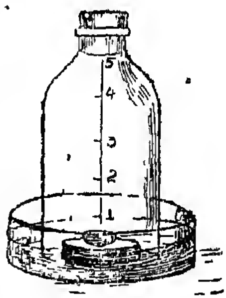


Fig 13.

Now light the phosphorus by touching it with the heated end of an iron rod. Quickly replace the air tight stopper. Phosphorus burns away into thick white fumes. Allow the jar to cool: the thick white fumes gradually disappear in water which rises inside the jar. Wait

till all the fumes inside the jar have dissolved in water; water rises further and fills about one-fifth of the space originally occupied by the air.

Remove the stopper again and lower a burning taper in the jar. It is at once put out, showing that the remaining gas is not the same as before.

EXPERIMENT 37. Place some iron filings in a muslin bag, and tie the bag to a piece of glass rod. Moisten the bag well, and place it in a jar inverted over water, (*See* Fig. 14). Examine the jar after a few days. It is found that the iron filings are covered with a reddish brown substance commonly known as rust. Water has risen in the jar showing that a portion of the air has been used by the iron in rusting. Raise the jar nearly to the surface of the water and place a glass disc tightly under the mouth of the jar so as to allow no water to escape. Put the jar upright and take into it a burning taper. Notice that the taper is at once put out. Measure the volume of water in the jar. Also find the whole volume of the jar. The fraction of the air that has been used up by the iron in rusting is found to be nearly one-fifth of the original volume of the air in the jar.

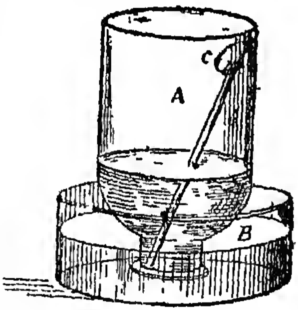


Fig. 14
A—inactive part of air.
B—water trough.
C—iron filings bag.

EXPERIMENT 38. Place some copper turnings in a hard glass tube open at both ends and fitted with

corks. One end is connected with a big Woulffe's bottle fitted with a thistle funnel and the other with a tube which dips in water under a beehive shelf placed in a trough.

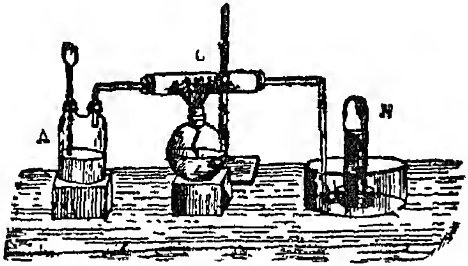


Fig 15

A—Woulfie's bottle. C—copper chips.
N—nitrogen.

(*See* Fig. 15.) A test tube full of water is inverted in the trough of water over the beehive shelf. Copper is strongly heated, and a current of air is made to pass over it by pouring water in the Woulffe's bottle. It is found that the copper turnings become black, when the air passes over them, because the active part of air—oxygen- combines with copper. The inactive part of air—nitrogen collects in the test tube; it does not combine with the copper. Remove the test tube when full, and take a burning taper in it. It is at once extinguished showing that the gas in the tube does not not help burning.

Moreover, if the volume of air which has come out of the Woulffe's bottle is measured and also the volume of air collected in the tube, it will be found that in passing over heated copper the air has last one-fifth of its volume.

From the above experiments we learn the following facts regarding air : —

1, Air is made up of two principal gases : the active part ; and the inactive part.

2. The active part of air known as oxygen, can combine with phosphorus, iron, heated copper, etc. and supports burning. The inactive part called nitrogen does not combine with any of the above substances and does not support burning.

3. Whenever the active part of air is removed, during a chemical change the volume of the air decreases by one-fifth.

4. About one-fifth of air is oxygen and four-fifth is nitrogen.

Air contains varying amounts of water vapour in it. This can be shown by the following experiments.

Experiment 39. Take a clear glass tumbler and dry it well from the outside. Put some pounded ice into the tumbler and allow it to stand for some time in air. Soon the outside surface of the glass tumbler is covered with small particles of water. The water vapour in the air, coming in contact with the cold surface of the tumbler, condenses into water drops.

Experiment 40. Place a few sticks of caustic soda in a flat dish and keep it exposed to air. Caustic soda has the property of absorbing moisture from air. After some time sticks of caustic soda become wet and if left for a long time dissolve in the water absorbed from air, and change into a liquid.

The following experiment proves the presence of carbon dioxide in air.

EXPERIMENT 41. Take some freshly prepared lime water in a flat dish and keep it exposed to air. After some time a white film of chalk (calcium carbonate) is formed on the surface of lime water. Now chalk can form on the surface of lime water only when it is acted upon by carbon dioxide. Hence it is clear that carbon dioxide has acted on it. So carbon dioxide must be in air for the lime water was exposed to air.

The air is not a chemical compound but a mechanical mixture The following reasons bear it out.

1. The composition of air varies though very slightly at different places and under different conditions, whereas every compound has always a definite composition.

2. The properties of a chemical compound are quite different from those of its constituents, whereas the properties of air lie midway between the properties of oxygen and nitrogen.

3. When oxygen and nitrogen are mixed together in the proportion in which they are present in air there is no production or absorption of heat and no change of volume takes place. But when a compound is formed there is some heat produced or absorbed and a change of volume takes place.

4. When air is shaken with water the dissolved air is richer in oxygen than the ordinary air; if air were a compound it would dissolve in water as a whole.

5. If air be liquefied and then allowed to evaporate nitrogen is given off first, leaving the remaining solution richer in oxygen. If air were a compound, this could not be the case.

Water of crystallization Many crystals when heated give off water and fall to powder. Blue vitriol or copper sulphate shows this peculiarity of some crystalline substances.

EXPERIMENT 42. Put a few crystals of copper sulphate in a dry test tube. Note their blue colour. Heat strongly. The crystals gradually lose their blue colour and become white powder. At the same time moisture deposits on the upper cool parts of the test tube showing that water is given off from the crystals.

To the white powder add a few drops of water and note the return of the blue colour.

If the white powder is dissolved again in water and the solution allowed to evaporate blue crystals may again be obtained. Water, thus, is necessary to form crystals of some substances. This water, therefore, which combines with a substance during crystallization, is known as the *water of crystallization*. Washing soda, alum, and borax contain a large proportion of such water.

Hygroscopic substances If calcium chloride or caustic soda be exposed to moist air, they take up moisture from the atmosphere. Such substances as take up moisture on being exposed to air are known as *hygroscopic* substances. Brown sugar (*gur*), quick-lime, sodium nitrate, etc., are other examples of hygroscopic substances.

Substances like caustic soda, calcium chloride, carbolic acid, potassium carbonate, etc., take up so much water on being exposed to air that they become liquids. Such substances are called *deliquescent* substances, and the phenomenon is called *deliquescence*.

Certain crystalline substances when exposed to air give up their water of crystallization and change into an amorphous powder, *e.g.*, sodium carbonate, hypo (sodium hyposulphate). Such substances are called *efflorescent*, and the phenomenon is known as *efflorescence*

SUMMARY

Composition of Air —Air is made up of the active part or oxygen and the inactive part, nitrogen, in the ratio 1 4 Other gases generally present in air are argon, carbon dioxide and water vapour

The active part of the air is removed by burning things in it

The air is not a chemical compound but a mechanical mixture of gases

Substances that take up moisture from the air are known as hygroscopic substances, *e.g*, caustic soda

If substances take up so much moisture on being exposed to air that they become liquids, then they are called deliquescent, *e g*, sodium sulphate anhydrous.

Efflorescent substances are those substances which give up their moisture to air when exposed to it and change into an amorphous powder, *e g*, sodium carbonates hydrated

QUESTIONS

1 Describe an experiment to show that air is made up of at least two gases. Give the properties of each gas

2 What conditions are necessary for rusting of iron? Describe experiments in support of your answer

3 Describe experiments to prove that water vapour and carbon dioxide are present in air How do you accout for their presence in air ?

4 Describe as fully as you can an experiment showing that air contains a gas in which a candle will not burn How would you try to find exactly the proportion of this gas in air ? (*P U 1914*)

5 Describe the composition of air What evidence would you give in proof of the statement that air is a mixture and not a chemical compound ?

6 Describe an experiment to show that air contains nitrogen equal to about four fifths of its volume

7 How would you show by experiments that only a portion of the air is consumed in (*a*) burning, (*b*) rusting ?

CHAPTER V.

Nitrogen.

Occurrence. Nitrogen is found in nature both in the free and the combined state. Nitrogen in the free state occurs in the air, of which it forms about four-fifths by volume. It is also present in small quantities in the volcanic gases. In the combined state it is widely distributed in nature in the form of nitrates. Nitrogen forms a necessary part of substances known as proteids which are present in plants and animal tissues.

Preparation. (1) Nitrogen can be prepared by the removal of oxygen from the air. As already studied in Chapter IV; oxygen can be removed from air by burning phosphorus, or allowing iron to rust in an enclosed air under a bell jar. Nitrogen obtained in any of these ways contains argon and other gases and thus is not quite pure.

(2) Pure nitrogen in the laboratory is obtained by heating ammonium nitrite, the chemical action being :—

Ammonium nitrite = Nitrogen + water.

Ammonium nitrite being a rare substance, in the laboratory a mixture of sodium nitrite and ammonium chloride is heated. The chemical action takes place in two stages as shown below:

Ammonium chloride + sodium nitrite

= Ammonium nitrite + sodium chloride.

Ammonium nitrite = Nitrogen + water.

EXPERIMENT 43. Take a round bottomed hard glass flask fitted with an air-tight cork through which passes a delivery tube. Fit the flask on a stand. Remove

the cork and put a little of the mixture of ammonium

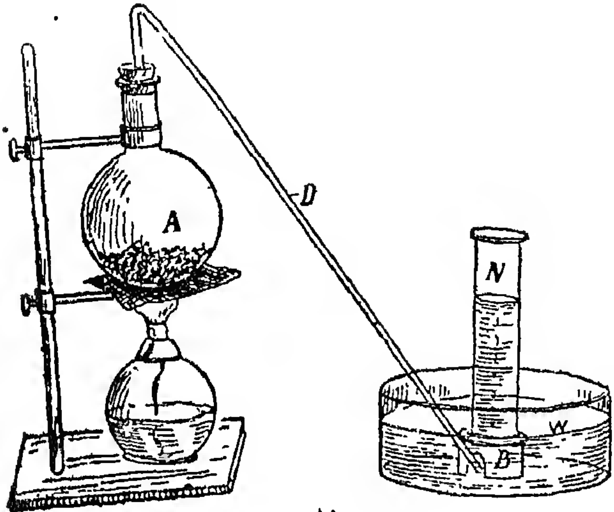


Fig 16.

A—mixture of ammonium chloride and sodium nitrite.
D—delivery tube. B—Beehive shelf. W—trough of water.
N—nitrogen

chloride and sodium nitrite in the flask. Add some water. Reinsert the cork. Dip the other end of the delivery tube in a trough of water under a beehive shelf. Heat the flask gently using wire gauze. Bubbles of gas begin to rise in water. Allow a few bubbles to escape and then collect the gas in jars over water as shown in Fig. 16. Perform the following experiments.

Experiment 44. Examine a jar full of the gas and note the gas has no colour. Smell and taste the gas. It has no smell or taste.

Put wet pieces of litmus paper, red and blue, into a jar full of the gas. Notice the gas has no

effect on either litmus paper. Nitrogen is neutral to litmus. Notice that water does not rise in the cylinder, showing that the gas is practically insoluble in water.

EXPERIMENT 45. Lower a lighted candle into a jar full of the gas; it at once goes out. Note also the gas does not burn itself.

Properties of Nitrogen :

Physical. Nitrogen is a colourless, odourless, and tasteless gas. It is very slightly lighter than air, it is very little soluble in water. It may be liquefied by strong cooling and increased pressure.

Chemical. It neither burns itself nor allows things to burn in it. It is neutral to litmus, and does not turn lime water milky. It is not poisonous, but animals die in it for want of oxygen. It is a very inert gas and combines directly with very few substances With magnesium, and calcium it combines directly at high temperature forming nitrides of magnesium and calcium :

Magnesium+nitrogen=Magnesium nitride.

Calcium+nitrogen=Calcium nitride.

Uses. Nitrogen in the air dilutes the oxygen, weakens its powers, and makes combustion much less intense than it would otherwise be. Atmospheric nitrogen is nowadays converted into its compounds with the help of electricity.

Oxygen

Occurrence Oxygen occurs as a gas free in the air to the extent of 23% by weight or 21% by volume. It occurs in combination in water, in almost all rocks, in plants, and in animals. It forms about 50% of the weight of the crust of the earth and is the most plentiful element known.

Preparation :—(1) Small quantities of oxygen can be prepared by heating strongly red oxide of mercury, when oxygen is given out and mercury left behind :

Mercury oxide = Mercury + oxygen.

EXPERIMENT 46. Put a small quantity of red oxide of mercury in a hard glass test-tube and heat over a flame. The red powder turns black and a mirror of mercury is formed on the cooler part of the tube. Take a glowing splinter of wood into the test tube, it bursts into flame.

Other oxides which give oxygen on heating are red lead, manganese dioxide, barium oxide, etc.

(2) Oxygen is also prepared by heating strongly potassium chlorate :

Potassium chlorate = Potassium chloride + oxygen.

If potassium chlorate is mixed with sand or black manganese dioxide, the oxygen from the chlorate comes off much rapidly and at a much lower temperature. This mixture is often spoken of as "oxygen mixture". The manganese dioxide takes no actual part in the chemical change. It helps the potassium chlorate to give off its oxygen more easily and more rapidly. Such a substance as makes a chemical action go much faster without itself undergoing any permanent chemical change at the end of the action is called a *catalytic agent*. Hence

(3) In the laboratory oxygen is obtained by heating the mixture of potassium chlorate and manganese dioxide (oxygen mixture). The chemical action that takes place is :

Potassium chlorate + manganese dioxide,

= Potassium chloride + oxygen + manganese dioxide.

EXPERIMENT 47. Fit up apparatus as in Fig. 17. It consists of a hard glass test tube fitted with an air tight cork through which passes a delivery tube whose other end dips in a trough of water under a beehive shelf. Fit the tube in a clamp stand. Remove the cork and mix about 10 grams of powdered potassium chlorate with about 3 grams of powdered manganese dioxide and put this oxygen mixture in the hard glass test tube. Reinsert the cork and heat gently

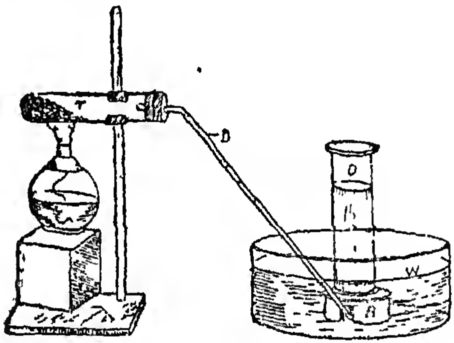


Fig 17.

T=hard-glass test tube. D—delivery tube.

B—beehive shelf. W—trough of water. O—oxygen.

beginning at the end nearer to to the cork and moving backwards. Bubbles of the gas begin to rise in water. Allow a few bubbles to escape and then collect over water, four or five jars of the gas. Keep each of them covered with a glass disc. At the end of the experiment remove the end of the delivery tube from under the water and then the burner; otherwise water will rush into the delivery tube and crack the test tube.

Let students find out the reason for this.

Manufacture of oxygen. (1) On a large scale for industrial purposes oxygen is obtained from the air. Air is liquefied by intense cold and great pressure in specially designed apparatus and is allowed to evaporate. Liquid nitrogen boils at a lower temperature than liquid oxygen Nitrogen, therefore, evaporates first, and the liquid left behind becomes richer and richer in oxygen. In this way fairly pure oxygen is obtained in large quantities.

Formerly barium monoxide (Baryta) was heated to dull redness in air; it took up oxygen to form barium peroxide which on further heating to bright redness gave up its oxygen which was collected for use. This method for the preparation of oxygen on a large scale is known as the Brin's process. This method is no longer in use now-a-days.

(2) Oxygen can also be obtained on a large scale from water by electrolysis (*See* Electrolysis of Water).

Take the jars in which oxygen has been collected and perform the following experiments:

Examine a gas jar and note the gas has no colour.

EXPERIMENT 48. Remove the disc and smell and taste the gas. It has no smell or taste.

Put moistened pieces of litmus papers, red and blue, into the same jar. Notice the gas has no effect on any of the litmus paper. Oxygen is neutral to litmus.

EXPERIMENT 49. Introduce a smouldering splinter into a jar full of the gas. The splinter re-kindles.

EXPERIMENT 50. Lower a burning candle into a jar full of the gas. It burns more brightly. Cover the jar with a disc. When the candle goes out, remove it, add a little freshly-prepared lime water in the jar and shake it. The lime water becomes milky, showing that carbon dioxide is produced. (Fig. 18.)

Fig 18.

EXPERIMENT 51. Put a little powdered sulphur in a deflagrating spoon and heat it over a flame. When sulphur begins to burn lower the spoon into a jar full of oxygen. The sulphur burns with a beautiful bluish flame.

EXPERIMENT 52. Repeat the above experiment with a piece of yellow phosphorus, in another jar. The phosphorus burns with a dazzling light and dense white fumes fill the jar.

Into the jars add a little of blue litmus solution; it turns red. Why?

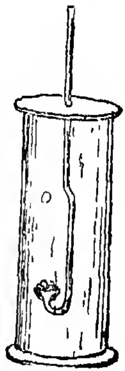


Fig 19
D. S. deflagrating spoon.

EXPERIMENT 53. Take a dry piece of sodium or potassium into a deflagrating spoon and heat it over

a flame. When it starts burning, lower it into a jar full of the gas. It burns with a bright yellow light.

Put in the jar red litmus solution; it turns blue. Why?

EXPERIMENT 54. Heat a bundle of steel wires or a fine coiled steel wire very strongly till it becomes red-hot, and lower it into a jar full of the gas. The wires burn giving out a shower of bright sparks.

Fig 20

S—coiled steel wire

O—oxygen.

Properties of Oxygen.

Physical. Oxygen has no colour, smell, or taste. It is very slightly heavier than air. The gas is only slightly soluble in water. This small amount of oxygen dissolved in water is sufficient for the respiration of fish. Oxygen has been liquefied and solidified under reduced temperature and high pressure. The liquid oxygen is pale blue in colour and the solid is white like snow.

Chemical. (*i*) Oxygen is neutral to litmus.

(*ii*) It does not burn itself but supports burning.

(*iii*) In oxygen things burn more brightly than in common air.

(*iv*) Substances like iron which do not burn in air, burn in oxygen.

(*v*) Substances like sulphur, phosphorus and charcoal burn in it forming substances which turn blue litmus solution red.

(*vi*) Substances like sodium, potassium and magnesium burn in oxygen to form compounds that turn red litmus solution blue.

(*vii*) It is very active and combines with most of the elements at high temperature forming compounds called oxides.

Test for Oxygen. Introduce a glowing splinter of wood into a jar full of the gas ; it bursts into a flame.

Uses of Oxygen. (*i*) Oxygen in the air supports animal and plant life, and is necessary for burning of substances.

(*ii*) Compressed oxygen is used for artificial respiration in hospitals, at high altitudes and in submarines.

(*iii*) Mixed with hydrogen it is used in the oxy-hydrogen flame for welding and cutting metals. If a compressed piece of quicklime is held in this flame, it gives an intensely white light known as the *lime-light.*

(*iv*) Mixed with acetylene it produces oxy-acetylene flame which is used in welding metals.

Reduction. The process of reduction is the reverse of oxidation. The removal of oxygen from any compound in part or totally is called reduction. Mercury oxide on heating is reduced to mercury, and oxygen gas is given out. Most of the oxides of metals such as zinc oxide, aluminium oxide, etc., when heated with coal under suitable conditions are reduced to the metal. Barium peroxide on heating alone is reduced to barium oxide and oxygen gas is given out. As we shall read later on hydrogen gas also reacts with some substances resulting in either the removal of oxygen from that substance or merely the addition of hydrogen to that element or compound. The action of hydrogen in this

way is also called reduction, *e.g.*, the combination of hydrogen with chlorine. The *reduction* is the removal of oxygen or the addition of hydrogen to any substance. Hydrogen, carbon and other substances that can bring about reduction are called *reducing agents*. Conversely oxidation is either the addition of oxygen to any element or compound or the removal of hydrogen from any compound; *e.g.*, when lead melts in air and is stirred it is oxidized to lead oxide, iron is oxidized to iron oxide (rust) in moist air. Similarly when sulphur burns in air or oxygen it is oxidized to sulphur dioxide.

Oxidation and Combustion The combination of a substance with oxygen is called *oxidation* and the substance is said to be oxidized. Like other chemical actions heat is produced in oxidation; but the production of heat is so slow that it diffuses away and little rise in temperature is noticed, as is the case with the rusting of iron.

When the combination of oxygen with a substance is rapid and is accompanied by both heat and light the process is called *combustion*. A substance that burns in a gas is called a *combustible* substance, and the gas in which the substance burns, is called a *supporter of combustion*. Thus substances like candle, sulphur, or charcoal that burn in oxygen or air are combustible, while oxygen or air is a supporter of combustion.

Some substances such as phosphorus oxidize slowly when exposed to air or oxygen without catching fire. Oily cotton rags, or damp hay or leaves if stored in heaps exposed to air, become heated to the ignition point and catch fire.

This slow process is known as ***auto-oxidation***. The fading of coloured fabrics on exposure to light is usually due to oxidation

The Oxides. When oxygen combines with an element, the substance formed is called an oxide

Thus :—

Sulphur+oxygen=Sulphur dioxide

Iron+oxygen=Iron oxide.

Sodium+oxygen=Oxide of sodium.

The oxides are divided into three classes :

1. The acidic oxides.
2. The basic oxides.
3. The neutral oxides.

The acidic oxides are those oxides which dissolve in water forming acids ; *e.g.*, oxides of sulphur, phosphorus, carbon, etc.

The oxides which when dissolved in water form bases are called the basic oxides ; *e.g.*, oxides of sodium, potassium, etc.

The oxides which have no action on litmus paper, are called the neutral oxides ; *e.g.*, oxide of hydrogen, carbon monoxide, etc.

SUMMARY

Nitrogen is prepared (*i*) from air by the removal of oxygen from it and (*ii*) by heating a mixture of sodium nitrite and ammonium chloride

Properties Colourless, odourless, tasteless, slightly lighter that air, a little soluble in water, neutral to litmus, incombustible, non supporter of combustion, inert It weakens the power of oxygen of the air

Oxygen is obtained by heating (*i*) red oxide of mercury or potassium chlorate, (*ii*) in the laboratory by heating a mixture of potassium chlorate and manganese dioxide The gas is collected over water

Properties. Colourless, odourless, tasteless, slightly heavier than air, only slightly soluble in water, neutral to litmus, incombustible, but supports combustion, very active, things burn in it more brightly

Uses Necessary to animals and plants, supports burning Used in artificial respiration

Oxidation is the union of elements with oxygen When oxidation is accompanied with heat and light, it is called combustion.

When oxygen combines with an element, the compound is called an oxide The oxides are (*i*) acidic, (*ii*) basic, (*iii*) neutral

QUESTIONS

1 How is nitrogen obtained from the air ? Sketch the apparatus used Is the nitrogen so obtained pure ?

2 How is pure nitrogen obtained in the laboratory ? Describe the chief properties and uses of the gas

3 What physical and chemical changes are observed when the following substances are heated in a test tube ?

(*a*) Red oxide of mercury
(*b*) Potassium chlorate
(c) Red oxide of lead

4. Describe the preparation of oxygen gas What experiment would you perform to show the properties of the gas ?

5 Explain the terms giving examples

(*a*) Catalytic agent

(*b*) Reduction.

(c) Oxidation

6 What is an oxide ? What are the chief kinds ? How do they differ from one another ? Give examples of each

7 What changes take place when a candle is burnt in a jar full of oxygen ? How would you show that these changes have taken place ?

8 Describe the properties of the two chief constituents of air In what proportion are they found ? A lighted taper burns more brilliantly in oxygen than in air Why ? (*P.U* 1920)

9 How is oxygen obtained on a large scale for industrial purposes ?

CHAPTER VI.

HYDROGEN AND WATER

Occurrence. Hydrogen occurs in the free state in volcanic gases, and in the outer atmosphere of the sun. It is the chief part of coal gas and water gas. Hydrogen is more commonly found in the combined state. With oxygen it forms water, with sulphur it forms sulphuretted hydrogen, and with carbon it forms hydrocarbons. It is an essential constituent of acids and alkalies, and all animal and vegetable tissues.

Preparation. (1) *By the decomposition of water by metals at the ordinary temperature.*

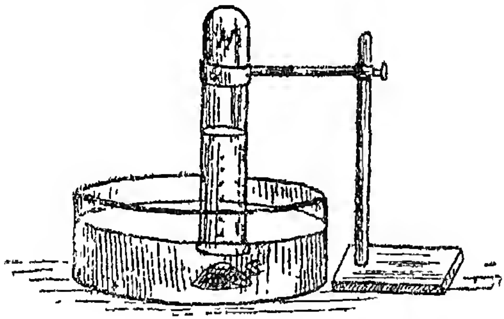


Fig 21

S—sodium in wire gauze.
H—hydrogen.

Experiment 55. Take a small piece of sodium and wrap it in a piece of wire gauze. Throw it in a trough of water and invert a test tube full of water over it. (*See* Fig. 21). As soon as sodium is throwny

in water, a hissing noise is produced and bubbles of gas rise and collect in the test tube by displacing the water.

Remove the test tube keeping its mouth downward, and apply a burning taper at its mouth. It will be found that the the gas burns with a noise.

EXPERIMENT 56. Rub a little sodium with an excess of mercury in a mortar; a substance known as sodium amalgam is formed. Take a little of sodium amalgam in a crucible and put it in a trough of water. Clap over it a wide test tube full of water as shown in Fig 22. Bubbles of gas rise up in the tube. When a sufficient quantity of the gas has collected in the test tube, remove it with its mouth downwards and bring a burning taper at its mouth. Notice the gas burns with a noise. Dip a red litmus paper in the solution in the trough; it is changed into blue. The solution feels soapy to tuch.

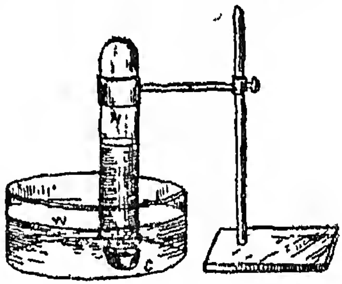


Fig 22

C—crucible containing sodium amalgam.
H—hydrogen. W—water.

In the above experiment, sodium reacts with water and hydrogen is given out, and caustic soda solution is formed in the trough;

Sodium + water = Sodium hydroxide + hydrogen.
(Caustic soda)

Like sodium, calcium and potassium also decompose water at the ordinary temperature. The action of potassium on water is so vigorous and so much heat

is given out that the hydrogen produced catches fire. Sodium and potassium are lighter than water but calcium sinks in water and the chemical action is comparatively slow.

(2) *By the decomposition of water by metals when heated.*

Some metals like iron, magnesium, zinc and aluminium decompose water when heated and hydrogen is obtained.

EXPERIMENT 57. Fit up the apparatus as shown in Fig. 23. Put some iron filings in the hard glass tube and heat them strongly. Boil water in a flask; and allow the steam to pass over the red-hot filings. Collect the gas in the jar as shown in the diagram. Test the gas in the jar as before; it will be found to be hydrogen.

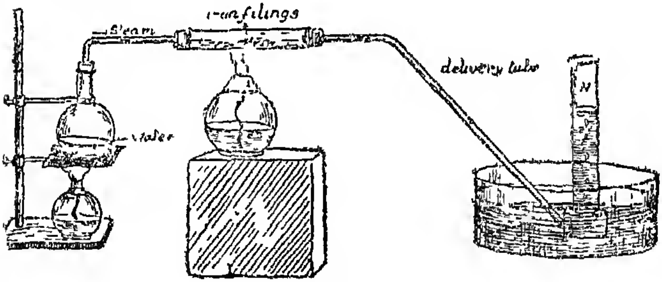


Fig. 23.

Red hot iron decomposes steam, combines with the oxygen forming iron oxide and sets free hydrogen:

Iron + steam = Iron oxide + hydrogen.

(3) *Decomposition of water by electricity.* (See *Electrolysis of Water*).

Water = Oxygen + Hydrogen.

(4) *From alkalies.* Certain metals like zinc and aluminium when warmed with solutions of caustic soda or caustic potash produce hydrogen :

Zinc + caustic soda = Sodium zincate + hydrogen.
Aluminium + caustic potash = Potassium aluminate + hydrogen.

(5) *From acids.* Hydrogen of acids is replaced by certain metals. Zinc, iron, magnesium, etc., when treated with dilute sulphuric or hydrochloric acid, set free hydrogen. Dilute nitric acid gives hydrogen with magnesium only.

In the laboratory, hydrogen is obtained by the action of dilute sulphuric acid on granulated zinc :

Zinc + dil. sulphuric acid = Zinc sulphate + hydrogen.

EXPERIMENT 58. Fit up the apparatus as shown in Fig. 24. In the Woulffe's bottle put granulated zinc. Fit a thistle funnel in one neck and a delivery tube in the other. Pour water in the bottle to cover the pieces of zinc. See that the lower end of the funnel dips under water and that the apparatus is air tight. Pass the other end of the delivery tube under a beehive shelf in a trough of water. Add some strong sulphuric acid through the thistle funnel. A brisk effervescence takes place and bubbles of hydrogen will be seen rising in water. Allow some gas to escape and then invert a cylinder

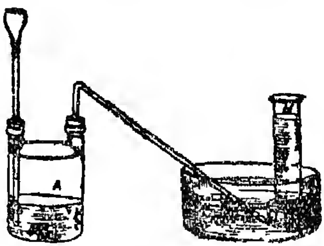


Fig 24

A—Woulff's bottle.
C—beehive shelf.
H—hydrogen.

full of water over the beehive shelf as shown in Fig. 24. The gas will collect in the cylinder. Fill a few cylinders with the gas and cover their mouths with discs and keep them inverted over the table.

Precautions.—Keep the apparatus air-tight and do not bring any flame near it, otherwise it may break with an explosion.

EXPERIMENT 59. Observe a jar full of the gas and note that it has got no colour. Remove the disc and smell the gas. It has got no smell and no taste. Put pieces of blue and red litmus paper in the jar. They remain unaffected. Hydrogen is neutral to litmus.

EXPERIMENT 60. (*a*) Keep a jar of the gas with mouth upwards and uncovered for some time. Bring a burning taper near the mouth. There is no explosion. This shows that there is no gas in the jar, it has escaped.

(*b*) Take a jar full of hydrogen and invert over it an empty cylinder as shown in diagram 25. Remove the upper jar after some time. Take a burning taper near its mouth keeping it inverted all the while. The gas burns with an explosion. This shows that the gas has gone from the lower to the upper cylinder. There is no gas in the cylinder which was formerly filled with hydrogen. Hydrogen, therefore, is lighter than air.

Fig 25

EXPERIMENT 61. Hold a cylinder full of the gas with its mouth downwards and introduce into it a burning candle as shown in diagram 26. Hydrogen burns at the mouth but the candle is extinguished. On taking out, the candle again starts burning on passing through the flame of burning hydrogen.

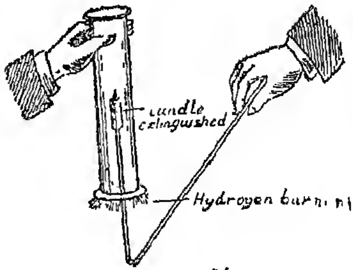


Fig 26

EXPERIMENT 62. Remove the delivery tube from the Woulffe's bottle and fit in its place a long drawn out tube as shown in Fig. 27. Pour more acid in the bottle and collect the gas by inverting a dry test tube over the jet. After a minute or two remove it and bring a burning match-stick near its mouth. The gas will burn with a noise. Repeat this twice or thrice till the gas burns quietly with a blue flame. Carry this flame of burning hydrogen to the jet from which the gas is escaping. The gas burns with a pale blue flame, which after a time becomes yellow on account of the glass becoming hot. Invert over the jet a wide tube open at both ends. Move it up and down. There will be a singing noise.

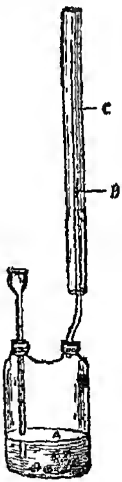


Fig. 27.
A—Woulffe's bottle with zinc and dil. sulphuric acid.
B—hydrogen burning.
C—wide tube.

EXPERIMENT 63. On the flame, place a clean beaker containing ice water. The beaker should be dried from outside. Notice the formation of drops of water on the outer surface of the beaker.

Pure hydrogen. The hydrogen preparad in the laboratory by Method 5 (Experiment 58) is not pure. Pure hydrogen is obtained in the laboratory by the action of dilute sulphuric acid on magnesium. The gas evolved is dried by passing it through a small flask containing strong sulphuric acid and is collected over mercury.

Properties of Hydrogen :

Physical. Hydrogen when pure has no colour, smell or taste. It is lighter than air, in fact it is the lightest gas known. It is only very slightly soluble in water.

Chemical. (*i*) It is neutral to litmus.

(*ii*) It burns with a blue flame but does not support combustion.

(*iii*) When mixed with air or oxygen, it forms an explosive mixture.

(*iv*) When it burns, it combines with oxygen to form water :

Hydrogen+oxygen=Water.

(Water is, therefore, an oxide of hydrogen or hydrogen oxide).

(*v*) It is a strong reducing agent.

Reduction. Hydrogen has a great tendency to unite with oxygen and is able to remove oxygen from

the compounds. Thus if hydrogen is passed over heated copper oxide contained in a hard glass tube it takes up oxygen from copper oxide and forms

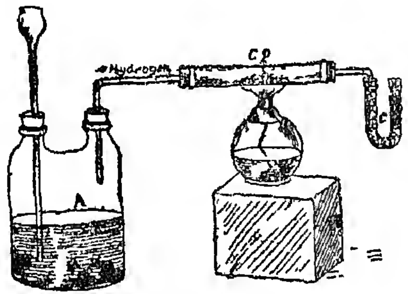


Fig. 28.

A—hydrogen generating apparatus.
CO —copper oxide.
C—calcium choride tube.

water. Copper oxide is changed into copper :

Copper oxide+hydrogen=Copper+water.

The water formed can be collected in tubes containing *calcium chloride* as shown in the diagram. (*See* Fig. 28.) Calcium chloride has the property of absorbing water vapour or moisture from gases. The removal of oxygen from a compound by its combination with hydrogen or some other substance like carbon or aluminium powder, is called *Reduction.* Substances like hydrogen, carbon, aluminium, etc., which remove oxygen from compounds are said to be *Reducing Agents.* The compound which loses oxygen is said to have been *reduced.*

Uses of Hydrogen. Hydrogen being the lightest gas known is used for filling balloons and airships but it is being replaced by coal gas and helium.

It is used in oxy-hydrogen flame for cutting and welding metals like platinum and iron and in the production of lime-light. It is chiefly used in converting liquid oils into solid fats.

Test If a burning taper is introduced in a jar full of hydrogen, the gas burns with a pale blue flame at the mouth but the taper goes out.

Water. Water is so well-known a substance that it hardly needs any description. It is very widely distributed in nature, nearly five-seventh of the surface of the earth is covered with water in oceans, seas, rivers, lakes, ponds, springs, etc. In the atmosphere it occurs in the form of water vapour. It forms a necessary part of all animal and vegetable substances. Next to air it is the most important thing for animal and vegetable life. The human body contains 70 per cent. of its weight of water.

Water is so plentiful in nature that it is unnecessary to prepare it in the laboratory. Its preparation simply involves the purification of a natural water by distillation.

Physical properties of water. When pure, water is a clear and tasteless liquid. Under ordinary circumstances it may be regarded as colourless but in reality it has a faintly bluish tinge which is seen when light is passed through a long tube of water closed at the ends with pieces of plate glass. (The deep blue colour of certain clear springs, like Panja Sahib at Hasan Abdal, Verinag in Kashmir, is due to light scattered from fine particles of solid matter in suspension.) Like most substances, water exists in three states; solid (ice), liquid (water) and vapour (steam). True

steam is invisible. The cloud of steam which issues out when water is boiled is a mist of small droplets of liquid water. Ice in large masses shows a distinct faint blue colour. Water freezes at 0°C. and boils at 100° C. Water expands when heated Its expansion is peculiar. From 0°—4°C, it contracts, beyond 4°C it expands. Thus at 4°C water has got maximum density, and then expands either on cooling or heating. It is on account of this property that water freezes only on the surface. The density of ice is less than of water, so that ice floats on water. 1 c.c. of water gives 1 09 c.c of ice . at the same temperature. Ice is a crystalline substance.

On account of its great solvent power water does not exist pure in nature. In order to get pure water for chemical purposes, it is distilled Distilled water has an unpleasant, 'flat' taste, and for drinking purposes it is aerated.

Natural waters. As already stated, water does not occur in a pure state. Even rain water which may be regarded as the purest form of natural water has got many gases dissolved in it besides the dust particles. The various kinds of natural water may be studied under the following headings : (*i*) Rain-water, (*ii*) River water, (*iii*) Spring or deep well water, (*iv*) Sea water, and (*v*) Mineral water.

The impurities in natural waters are of two kinds : (1) Suspended impurities, both mineral and organic ; (2) Dissolved impurities, both solids (mineral and organic), and gases. Organic impurities in water are due to the decay and growth of animal and vegetable matter.

The amount and nature of impurities depend chiefly on the source of water.

Rain water. Water from seas, rivers, lakes, etc., is being constantly evaporated by the heat of the sun. The vapour rises up in higher and colder regions, becomes condensed and falls in the form of rain. It is thus a process of distillation in nature on a very large scale. Rain water, therefore, should be pure. It, however, in its journey through the atmosphere dissolves the gaseous impurities in the air, *e.g.*, ammonia, carbon dioxide, nitric acid, etc. Rain water becomes a bit purer after the first shower. Rain water collected in country-side is purer than that collected in town. Besides gaseous impurities it has got suspended in it insoluble dust particles taken from the atmosphere.

River water. Rain water in the form of streams is carried to rivers. As water moves on the surface of the soil it takes up salts and organic matter from the cultivated areas or decaying vegetable matter or from sewages from cities which are thrown into rivers, and suspended matter such as clay. The nature of salts present in water, therefore, depends upon the soil through which it passes. Thus, if a river passes, through a limestone district, it contains calcium carbonate dissolved in it ; if it passes through a salt district, it contains salt, and so on.

Spring or deep well water. The only difference between river water and spring water is that the latter does not contain suspended impurities as it undergoes a sort of filtration while passing through porous beds. The organic impurities also decrease but mineral impurities increase as water in its course through the porous layers dissolves many salts. Thus, the nature of impurities depends upon the soil through which the water passes. This water is quite good for drinking purposes.

Sea water. Rivers carry their impurities both suspended and dissolved to the sea. Evaporation is constantly going on, thus increasing the precentage of dissolved impurities. Sea water is the most impure form of all natural waters. The chief impurities are sodium chloride (2·6%) and bromides, sulphates and carbonates of calcium, magnesium and potassium.

Mineral water When spring water contains large amount of dissolved substance so as to give to it a marked taste, it is known as *mineral water*. Such water possesses often some medicinal properties. Such springs are common in India. Mineral springs are named after the important substance dissolved in them. Thus, sulphur springs contain sulphuretted hydrogen dissolved in water. There are such springs at Anantnag (Kashmir), Namaldam (Mianwali), Jwalamukhi (Kangra) and Nathiagali (Abbottabad). These springs have the property of curing skin diseases.

The springs at Jwalamukhi contain bromide and iodide of sodium and potassium. There are *chalybeate* waters which contain iron carbonate dissolved in them. Again, there are some springs which contain carbon dioxide disscived in a large quantity—*effervescent springs*. Epsom salts are present in certain springs and they are used as purging places, *e.g.*, Grot in Jhelum district. Such a water is known as *bitter water*.

Drinking water. It is of the hightest importance that drinking water should be clear and colourless, and as free as possible from organic impurities. Many dissolved impurities, *e.g.*, some mineral salts are quite harmless. Water becomes foul when it contains organic impurities and harmful bacteria which may be the cause of cholera and typhoid diseases. Drinking

water in cities, as already mentioned, is made harmless by filtering it through beds of sand and gravel. It is exposed to air and sunlight to kill germs, and in some places the germs are destroyed by adding chlorine or bleaching powder solution. River water, especially When it passes near a city, should not be used for drinking purposes. The source of water supply in a city should be kept free from contamination.

Hard and Soft water When we wash our hands or clothes in rain water collected in tanks with soap, we find that a lather is produced easily If we use river water, we notice that lather is not formed so easily and that more soap is required to produce lather. Water in which soap lathers easily is called *Soft water* and the other in which lather is produced with difficulty is termed as *Hard water*. In the case of hard water, soap produces a curd-like precipitate which floats on the surface.

Causes of hardness If we use rain water or distilled water for washing purposes we see that soap lathers readily, but if we use river water we see that much soap is to be washed before we get good lather. Rain water or distilled water does not contain impurities dissolved in it, but river water contains a large amount of dissolved impurities. Hardness, therefore, must be due to the presence of dissolved impurities in water. Water has generally got carbon dioxide dissolved in it in the form of carbonic acid. The presence of carbonic acid in water imparts to it the property of dissolving chalk (calcium carbonate) to a large extent. The calcium carbonate, however, is not present as such in water but is present as a bicarbonate;

Water+calcium carbonate+carbon dioxide
=Calcium bicarbonate+water.

It is the presence of this salt and other salts of calcium and magnesium which renders water hard. The extent of hardness differs with different kinds of water. Sea water is the hardest of all, next comes river or spring water and rain water is usually soft.

EXPERIMENT 64. Prepare a soap solution by heating slices of pure soap in distilled water. Take three stoppered bottles, one containing 20 c.c. of distilled water, the other 30 c.c of tap water and the third 20 c.c. of water in which magnesium sulphate has been dissolved. Add to each 20 c.c. of soap solution and shake well and allow to stand. Notice that lather is formed very readily in distilled water, not so readily in tap water and that curd-like specks are formed in the third case Put more soap solution in the third bottle and shake. Now you will get a lather.

It will be seen from the above experiment that distilled water requires the least quantity of soap solution while water containing magnesia in it the greatest. Hence this water is very hard.

EXPERIMENT 65. Pass carbon dioxide in a beaker containing freshly prepared lime-water. Notice that lime-water becomes milky. Continue passing carbon dioxide in it for a long time till the milkiness disappear and the water again becomes clear. Take 20 c.c. of the liquid in a beaker and find out the quantity of soap solution required to form permanent lather. Boil the remaining liquid and see that it again becomes milky. Allow it to stand and cool. Then take 20 c.c. of this cooled liquid in a beaker and find out as above the quantity of soap solution required to produce permanent lather. In another 20 c.c. of the clear liquid add a little lime water and then find out the quantity of soap solution required to produc

permanent lather. From these experiments it will be seen that more soap solution will be required in the first than in the second and third cases.

EXPERIMENT 66. Take some water which contains magnesium sulphate dissolved in it. Boil the solution and allow it to cool. Take 20 c.c. of this solution and find out the amount of soap solution required to produce lather. It requires a large amount of soap solution to form permanent lather.

As already stated, water containing carbon dioxide dissolved in it is able to atttack calcium carbonate and holds it in the form of calcium bicarbonate. The change is the same as in the case of Experiment No. 65 when we pass carbon dioxide in lime water for a long time. The lime water turns milky on account of the formation of calcium carbonate which being insoluble in water remains suspended in it and imparts to it a milky appearance. When more carbon dioxide is passed through this water, calcium carbonate combines, with carbon dioxide and forms calcium bicarbonate :

Calcium oxide in solution + carbon dioxide
(Lime-water)
= Calcium carbonate + water

Calcium carbonate + carbon dioxide + water = Calcium bicarbonate + water

It is the calcium bicarbonate in water which makes water hard. We have seen in the above experiment that if such water is boiled, less quantity of soap solution is required to produce permanent lather or it becomes soft. The reason is that when this water is heated calcium bicarbonete is decomposed, gives up carbon dioxide anc

forms calcium carbonate. Calcium carbonate is insoluble in water and settles down, thus making the water soft:

$$\text{Calcium bicarbonate} \xrightarrow{\text{heat}} \text{Calcium carbonate (insoluble)} + \text{carbon dioxide} + \text{water.}$$

In Experiment No. 66 we find that even after boiling curd-like precipitate is produced or, in other words, the water is not rendered soft. Water which on boiling becomes soft is known as *Temporary Hard* water and the other which does not lose its hardness on boiling is termed as *Permanent Hard* water. Temporary hardness of water is thus due to calcium, magnesium and iron bicarbonate in water. Permanent hardness is due to the sulphates and chlorides of calcium and magnesium.

Softening of hard water. We have seen above that temporary hard water can be softened by boiling it when carbon dioxide is driven out. In the Experiment No. 65 while taking 20 c.c. of liquid in which we add some lime-water, we find that a small quantity of soap produces lather or in other words the addition of lime-water softens the water. Thus another method of softening temporary hard water is to add an *exact* quantity of lime-water to it. The lime reacts with calcium bicarbonate and produces insoluble calcium carbonate which settles at the bottom and the water is rendered soft.

Calcium bicarbonate + calcium oxide (in water) = Calcium carbonate (insoluble) + water.

Similarly magnesium bicarbonate is removed but by a *larger* quantity of lime.

Permanent hardness, as already mentioned, is due to compounds of calcium and magnesium (chlorides

and sulphates). These are soluble in water and hence cannot be removed by boiling. Lime also has no effect. In order to remove permanent hardness, washing soda is added:

Calcium sulphate or chloride + sodium carbonate (washing soda)
= Calcium carbonate (insoluble) + sodium sulphate or chloride soluble

Magnesium sulphate or chloride + sodium carbonate
= Magnesium carbonate (insoluble) + sodium sulphate or chloride (soluble)

The insoluble carbonates settle at the bottom and the water becomes soft; whereas sodium salts have no effect on soap solutions.

The method usually employed is distillation. Distilled water, as already stated, is soft.

In many cases both calcium or magnesium bicarbonates and their chlorides or sulphates are present in water. Such water will only be partially softened by boiling and the boiled water will require an addition of washing soda to soften it completely. Again sodium chloride in water makes the production of lather difficult and requires a large quantity of soap to be wasted, because soap is not soluble in salt water as is the case with sea water. A special kind of soap known as marine soap is used for washing purposes with the sea water.

Consequences of hardness of water. When hard water is used in a steam boiler, salts deposit on the tubes of the boiler as 'boiler crust' or 'fur' which does not allow heat from the metal to reach the water readily. The life of the boiler is shortened. If the crust is not removed, the boiler may explode. Hard water used for washing in the

household or laundry, results in the waste of a large quantity of soap because much of the soap is first used up to remove the salts in solution in the form of precipitates. It is after the romoval of salts only that a lather will be produced and the clearing effect of soap begins.

Chemical properties of water. It has been already seen that water combines with the oxides of metals such as those of sodium and potassium to form bases; and with the oxides of non-metals such as sulphur, phosphorus, etc., to form acids.

Water has the property of combining directly with many salts Such salts crystallize out from their aqueous solutions with a definite proportion of water known as the *water of crystallization.* Thus crystals of alum, copper sulphate, borax, nitre, always contain a definite proportion of water of crystallization.

In the preparation of hydrogen we have studied that the action of metals on water differs with the nature of metals. Metals, like sodium, potassium and calcium, decompose it at the ordinary temperature; iron, zinc, magnesium, aluminium do so only when heated and gold and silver do not decompose it at all. In the case of decomposition, the oxide of the metal is produced and hydrogen is set free.

Composition of water. Water is composed of oxygen and hydrogen in the ratio of 1:2 by volume and 8:1 by weight. The composition of water by volume can be determined by (i) analysis of water, *i.e.*, breaking it up into its constituents, and (ii) synthesis *i.e.*, preparing water by combining oxygen and hydrogen.

Analysis of water.

EXPERIMENT 67. In a *voltameter* put some acidulated water Invert two test tubes of the same size and shape full of acidulated water over the electrodes. Connect the binding screws with the poles of a battery with copper wires. Bubbles of gas begin to rise from each electrode and collect in the test tubes. Allow the current to pass for about half an hour. You will find that the volume of gas in the tube above the kathode is double the volume of gas in the tube over the anode. Stop the current and remove the test-tubes. Test the

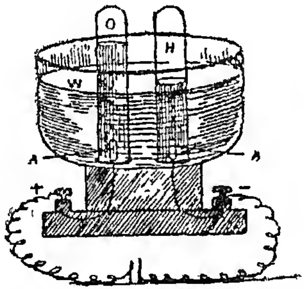


Fig 30.

A.—anode, O—oxygen
K—kathode. H—hydrogen.

gases. The tube over the kathode contains hydrogen and the other oxygen. Sulphuric acid is added to make the water a conductor of the current. Pure water is a non-conductor.

Synthesis of water. The composition of water can be more accurately determined by taking hydrogen

and oxygen in the ratio of 2:1 in a *Eudiometer tube* and making them combine by passing an electric spark in the mixture.

EXPERIMENT 68. Take a *Eudiometer tube*— a graduated glass tube closed at one end and having two short platinum wires fused near the closed end (*See* Fig. 30.) Fill it completely with mercury, close its mouth with your thumb and invert it carefully in a trough of mercury. Pass in to the eudiometer tube oxygen and hydrogen in the ratio of 1 : 2. Now connect the platinum wires with the poles of a battery. Sparks are produced and an explosion occurs. After a time mercury rises up and completely fills the eudiometer tube the volume of water formed being very small as compared with that of the original gases. Droplets of water are seen on the walls of the tube after cooling This experiment clearly shows that water is made up of two volumes of hydrogen and one volume of oxygen. If any of the gases is used in excess, it is left behind.

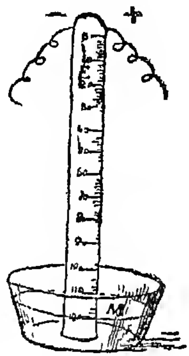


Fig. 30
M—mercury.

Composition of water by weight The composition of water by weight can be determined by passing pure and dry hydrogen over a weighed amount of heated copper oxide Hydrogen reduces copper oxide into copper and forms water which is collected and weighed.

EXPERIMENT 69. Arrange the apparatus as shown in the Fig. 31. A is the apparatus for the preparation of hydrogen better by the action of magnesium on dilute sulphuric acid. Hydrogen is passed through B containing strong sulphuric acid. The dry hydrogen is then allowed to pass through *c*, a hard glass tube containing copper oxide. Weigh the tube *c*, with its connecting tubes and copper oxide. Weigh the U-shaped tube *d*, containing calcium chloride to absorb the water formed. Connect the apparatus. Pass hydrogen to drive out air from the apparatus. Then heat the tube containing copper oxide from end to end so that it may not break. Before heating see that the apparatus is air-tight. Continue heating till the whole copper oxide is reduced to copper. Stop heating but continue passing hydrogen till the tube is cooled. Disconnect the apparatus. Weigh the tube *c* with its contents and connecting tubes. Then weigh the U-shaped tube *d*. Calculate as follows :—

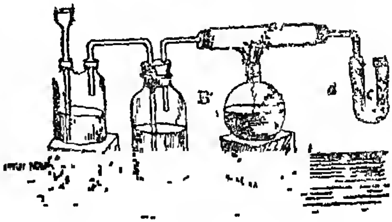


Fig 31.

d—calcium chloride tube.
A—hydrogen apparatus.
B—strong sulphuric acid.
c—copper oxide.

Weight of hard glass tube (*c*) with connecting tubes and copper oxide	= 45	gm.
Weight of U-shaped tube (*d*) with its contents	= 80	gm.
Weight of (*c*) after the experiment	= 42·6	gm.

$\therefore$ Weight of oxygen lost $= 45 - 42{\cdot}6 \qquad = 2{\cdot}4$ gm,

Weight of U-tube after experiment $\qquad = 82{\cdot}7$ gm.

$\therefore$ Weight of water formed $= 82 \cdot 7 - 80 = 2 \cdot 7$ gm.

$\therefore$ Weight of hydrogen $= 2.7 - 2 \cdot 4 = .3$ gm.

Hence the ratio between oxygen and hydrogen

$= 2 \cdot 4 : \cdot 3$ or $24 : 3$ or $8 : 1$.

SUMMARY.

Hydrogen is prepared in the laboratory by the action of dilute sulphuric acid on zinc

Properties Hydrogen is a colourless, tasteless, odourless, lightest known gas and very slightly soluble in water It is neutral to litmus It burns with a blue flame but does not support combustion When it burns in air or oxygen water is produced It is a reducing agent It forms an explosive mixture when mixed with air or oxygen It is used in oxyhydrogen flame for fusing and welding metals It is used in filling balloons and air-ships and for solidifying fats

Test It burns with a blue flame but does not allow the taper to burn in it

Water Water is found in nature as rain water, sea water, river water, spring water and mineral water Rain water is the purest form and sea water the most impure form of water in nature

Natural water contains both suspended and soluble impurities

Soluble impurities such as bicarbonate, chlorides and sulphates of calcium and magnesium render water hard

Hard water is that in which soap lathers with difficulty.

Soft water is that in which soap lathers easily

Hardness of water is either temporary or permanent.

Temporary hardness is removed by (*i*)boiling, (*ii*)by adding lime.

Permanent hardness is removed by (*i*)washing soda, or (*ii*) distillation.

Potassium, sodium, and calcium liberate hydrogen from cold water Magnesium, zinc and iron liberate hydrogen from steam Gold and silver have no action on either cold or hot water

Two volumes of hydrogen unite with one volume of oxygen to form water Eight parts of oxygen by weight unite with one part of hydrogen by weight to form water Water is decomposed into its elements by passing electric current through it It is formed by passing electric sparks in a mixture of two volumes of hydrogen and one volume of oxygen

QUESTIONS

1. Describe the preparation and chief properties of hydrogen Sketch the apparatus State the chief precautions to be observed in the preparation of this gas

2 Hydrogen is said to be an element, water a compound What practical proof can you give of the truth of these statements

3 Describe an experiment to show that water is an oxide of hydrogen,

4 A number of cylinders are filled with a gas, which may be either hydrogen or nitrogen How would you find experimentally which gas is present ?

5 Describe two different methods by which hydrogen may be obtained from water

6 Hydrogen is passed over strongly heated, copper oxide Describe the changes which take place and name the substances produced

7 What do you understand by the terms analysis, synthesis ? Explain how the composition of water may be shown (*i*) by analysis, (*ii*) by synthesis

8 Describe what occurs when steam is passed over red-hot iron Sketch and describe the apparatus used

9 How will you show that air contains one of the consituents of water

10 What is a voltameter ? Explain how it is used to determine the composition of water by volume

11 Explain what is meant by dissolved and suspended impurities in natual water How can these be removed ?

12 In what respects do rain water, river water and sea water differ from one another ? How would you prepare fresh water from a dirty sample of sea water ?

13 You are supplied with some muddy river water How would you proceed to purify it ? Sketch the apparatus you would use

14 Write a careful account of the action of heated magnesium and iron on steam

15 Define the terms 'hard' and 'soft' in relation to water, Distinguish between temporary an permanent hardness, state and explain the action of soap on hard water, How may permanently hard water be softened ?

16 Write a brief account of natural waters

17 Distinguish between temporary and permanent hardness of water Name the substances to which each is respectively due How can temporary and permanent hardness be removed from water ?

18 Hard water is (*a*) boiled, (*b*) mixed with lime-water, (*c*) mixed with washing soda Describe and explain what happens in each case

CHAPTER VII.

SULPHUR.

Occurrence Sulphur occurs in nature both in the free and in the combined state. Free or native sulphur is abundant in volcanic districts in Sicily, America, Japan and China. In India large deposits are found at Sanni, in Kelat State. Small quantities are found on the volcanoes of Koh-i-Sultan in Baluchistan and Jwalamukhi in the Punjab. The crude sulphur is very impure and contains as impurities pumice stone and other minerals.

In the combined state it occurs in the form of metallic sulphides and sulphates. Combined with hydrogen, it is found as *sulphuretted hydrogen*. Sulphide is a compound of a metal with sulphur Some common sulphides met with in nature are (1) Galena—lead sulphide, (2) zinc blend—zinc sulphide; (3) iron pyrites—iron sulphide and (3) copper pyrites—copper sulphide Their compounds are the ores of the metals; *i.e.*, metals are extracted from them easily and cheaply.

The common sulphates, *i e*, compounds of sulphur, oxygen and a metal are,—(1) Gypsum—calcium sulphate; (2) heavy spar or barium sulphate and (3) magnesia (Epsom salts)—magnesium sulphate. Gypsum is very common in the Punjab in the Salt Range.

Sulphur is a constituent of eggs, onions, garlic, mustard' etc , and also hair and wool.

Extraction. Native sulphur is always more or less mixed with earthy and mineral impurities, from which it is necessary to free it. The process is carried out as follows:—

Native sulphur is piled up in a brick kiln on a sloping floor, vertical air spaces being left at intervals. They are covered with powder ore. The ore is kindled at the top, and the heat produced by the

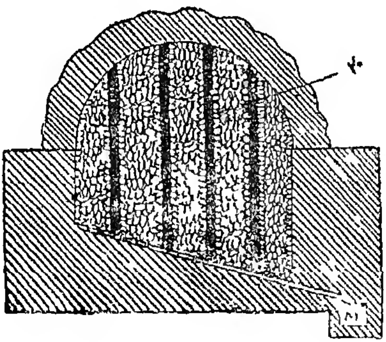


Fig 32

K—kiln

M—molten sulphur

burning of a part of sulphur melts out the remainder. The melted sulphur runs down the sloping floor and is collected into wooden moulds. This is a wastefull process, as much of the sulphur is burned away.

The sulphur obtained by the above process is further refined by distillation in iron retorts. The vapours are passed into a brick chamber. At first the vapour condenses on the old walls as a light yellow crystalline powder called Flowers of Sulphur. As the walls become hot this melts (unless it is removed) and runs down as a liquid to the bottom, from where it is run off into moulds and cast into sticks known familiarly as 'brim stone'. or roll sulphur'.

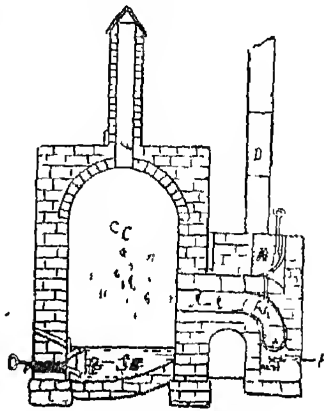


Fig 33
F—furnace E—exit
R—retort D—chimney
C—cold chamber
S—sulphur

Sulphur is also prepared by roasting iron pyrites in kilns where iron oxide and sulphur and sulphur dioxide are formed.

Iron pyrites = Iron oxide + sulphur dioxide + sulphur

Physical properties of sulphur. Sulphur is a pale yellow, brittle, crystalline solid. It has got a faint odour but no taste. It is insoluble in water but is soluble in carbon disulphide. Its density is 2·06. It is a very bad conductor of heat and electricity and when rubbed with flannel or catskin becomes charged. It melts at 114°C., into a thin, pale yellow, amber-coloured liquid. When heated further, its colour darkens and it becomes thicker and thicker till at about 180° C., it becomes almost black and so thick

that it can no longer be poured from the vessel. On further heating the liquid becomes thin but remains black in colour. At 444° C. the liquid boils and is converted into dark orange vapour. On cooling, the above changes are reversed

EXPERIMENT 70: Heat some powdered sulphur in a clean dry test tube and study the changes as described above.

The Allotropic Modification of sulphur. Sulphur occurs, mainly in three different forms, two crystalline and one amorphous. The crystalline varieties are rhombic or octahedral and the monoclinic or prismatic sulphur; the amorphous is known as the plastic sulphur. This property possessed by certain elements of existing in more than one form with different physical properties is known as *allotropy* and the less common forms are said to be the *allotropic modifications or allotropes* of the more common form. We shall come across other examples of elements showing allotropy later on.

Rhombic or Octahedral Sulphur.

Preparation. EXPERIMENT 71. Take some powdered roll sulphur in a dish and add carbon disulphide to it. The sulphur dissolves. Filter the solution in a clean dish and allow the filtrate to evaporate. Pale yellow octahedral crystals will be formed. This is rhombic sulphur. Don't bring a flame near carbon disulphide but simply keep it exposed to air.

Properties, It is a pale yellow crystalline substance with density 2·03 to 2·06 and melting point 114°. It is insoluble in water but is soluble in carbon disulphide.

It is the stable form at the ordinary temperature and all other forms pass into it on standing. Roll sulphur and flowers of sulphur contain this variety.

Prismatic or Monoclinic sulphur

Preparation EXPERIMENT 72. In a dish put some powdered roll sulphur and heat it over a sand bath. When the whole of it has just melted allow it to cool. A crust will be formed on the surface. Make two holes in the crust with a pointed glass rod and pour off the liquid. Remove the crust carefully. You will find inside the dish fine, transparent, needle-shaped crystals of prismatic sulphur.

Properties. Its density is 1 · 96 and melts at 19_0C. It is insoluble in water but soluble in carbon bisulphide. On evaporating the solution we get rhombic sulphur. Rhombic sulphur when heated above 90°C. gives this variety.

Plastic Sulphur.

Preparation. EPERIMENT 73. Take some powdered roll sulphur in a dish and heat it, it melts into a liquid; heat it further till it is converted into a thick dark coloured viscous liquid. Heat it still further till the liquid again becomes thin and is about to boil. (The temperature should be about 35°C.) Pour this liquid in a trough full of cold water. You will get a soft, rubber-like substance known as Plastic Sulphur. It can be pulled out like rubber and is as elastic.

Properties Its density is 1 ·95. It is insoluble in carbon disulphide. On standing for a few days, it becomes opaque, hard and brittle and is partly converted into rhombic sulphur.

Chemical properties of sulphur. The three varieties though they differ physically are chemically identical. When a given weight of any of the varieties is burned in air or oxygen, we get an equal quantity of sulphur dioxide in each case and nothing is left behind.

Sulphur + oxygen = Sulphur dioxide.

Sulphur combines with many metals like iron, copper, zinc directly forming sulphides :

Iron + sulphur = Iron sulphide.
Copper + sulphur = Copper sulphide

Sulphur combines directly with chlorine to form an amber-coloured liquid called sulphur monochloride, which is used in vulcanizing rubber. With carbon sulphur gives carbon bisulphide.

Uses of sulphur Crude sulphur is used for making sulphur dioxide, for sulphuric acid and the manufactures of carbon disulphide. Refined sulphur is used in medicine and in the preparation of gunpowder, matches, fireworks and dyes. It is used in large quantities for vulcanizing rubber. It is used as a disinfectant by burning it in the room to be disinfected. It is used for dusting vines to prevent the growth of fungus.

Componnds of Sulphur. The most important compounds are its oxides—sulphur dioxide and sulphur trioxide.

Sulphur dioxide *Occurrence.* It occurs in the volcanic gases and in small traces in the atmosphere of cities where it is obtained by the burning of sulphur present in coal.

Preparation. (1) It can be prepared by burning sulphur in air or oxygon :

Sulphur+oxygen=Sulphur dioxide.

(2) By roasting iron pyrites in a kiln :

Iron pyrites=Iron oxide+sulphur dioxide
+sulphur.

(3) By the action of concentrated sulphuric acid on copper, mercury, zinc, sulphur, charcoal, etc. In the laboratory we prepare it by heating strong sulphuric acid and copper. The chemical change is

Copper+sulphuric acid=Copper sulphate
+sulphur dioxide+water.

EXPERIMENT 74. Fit up the apparatus as shown in Fig. 34. In the flask put copper turnings. Pour sufficient quantity of strong sulphuric acid through the thistle funnel. See that the lower end of the thistle funnel dips under the acid. Heat it gently on a wire gauze. The mixture in the flask will become dark in colour and a gas will be given off. Collect the gas in jars, by downward displacement of air as shown in Fig. 34. In order to see whether the cylinder is full of gas or not, take a burning taper near the mouth If it goes out, the jar is full. Collect a few jars and perform with them the following experiments —

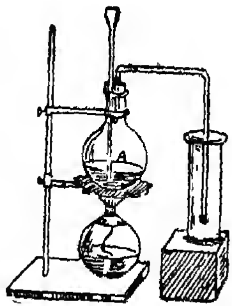


Fig. 34.
A—flask containing copper and strong sulphuric acid.

EXPERIMENT 75. Examine the colour of the gas in a cylinder, smell the gas. Put a moistened blue

litmus paper in the jar. See that it is at once turned into red.

EXPERIMENT 76 Invert a jar of gas on another jar and test the presence of gas in the two cylinders after some time. The burning taper will be extinguished in the lower jar which was empty before and continues to burn in the upper which was full. This shows that the gas can be poured downwards from one cylinder into the other *i.e.*, it is heavier than air. (*See* Fig. 35.)

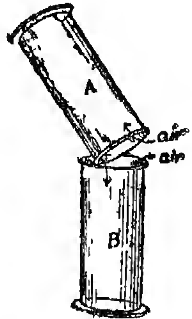


Fig 35.

EXPERIMENT 77. Invert a cylinder full of the gas in a trough of water. Water will rise in the cylinder. Place a disc at the mouth of the cylinder and remove it from the trough and place it on the table with mouth upwards. Throw a blue litmus paper in the liquid in the cylinder. The colour is at once changed into red. Taste the liquid. It is slightly sour.

EXPERIMENT 78. Introduce a burning taper in a cylinder full of the gas. It at once goes out. The gas does not burn.

EXPERIMENT 79 Prepare a weak solution of potassium permanganate in a tube and add to it a drop of sulphuric acid. Pass sulphur dioxide in the tube. Notice that the pink colour disappears.

EXPERIMENT 80. Prepare a solution of potassium dichromate and add to it a drop of sulphuric acid.

Soak a piece of filter paper in it and throw the piece in a cylinder full of the gas. The deep yellow colour will change into green. This is used as a test for the gas.

EXPERIMENT 81 Moisten a few coloured flowers and throw them in a jar full of the gas. The colour is destroyed and the flowers are bleached. Add a few drops of sulphuric acid to the bleached flowers. The colour is again restored. Repeat the same experiment with writing-ink ; the colour fades or is bleached.

Bleaching action of sulphur dioxide or sulphurous acid. From moisture they take up oxygen and set free hydrogen which being in the nascent (newly) born) condition is a powerful reducing agent. The hydrogen reduces the colouring matter by combining with it and changing it into a colourless compound.

Properties :- Physical. Sulphur dioxide is a colourless gas with a strong choking smell like that of burning sulphur. It is poisonous and attacks the eyes. It is 2·3 times heavier than air. It is freely soluble in water forming an acid solution—sulphurous acid :

Sulphur dioxide+water=Sulphurous acid.

It can be liquefied, the liquid on cooling produces cold.

Chemical. (1) It does not burn nor allow any substance to burn in it, (2) It is acidic in property—turns blue litmus red. (3) When dissolved in water it forms sulphurous acid which has got all the properties of the gas. (4) It is a reducing agent—destroys the colour of potassium permanganate or dichromate. (5) Has got no action on lime-water. (6) It is a bleaching agent—bleaches vegetable colours which can be restored on exposure to air or adding a few drops of sulphuric acid. (7) The solution of the gas in water if left exposed takes up oxygen from the air and forms sulphuric acid.

Sulphurous acid+oxygen=Sulphuric acid.

The salts of sulphurous acid are known sulphites:

Sodium hydroxide+sulphurous acid=Sodium sulphite+water.

Uses. Sulphur dioxide is employed in the manufacture of sulphuric acid. It is used as a bleaching agent for delicate fabrics like straw, silk and wool. It is used as a germicide It is also used in the manufacture of ice

Test. (*i*) Filter paper dipped in acidified solution of potasuium dichromate is turned from yellow to green in colour. (*ii*) Its characteristic smell. (*iii*) It bleaches vegetable colours.

Sulphur Trioxide. *Preparation.* It is usually prepared by passing a dry mixture of sulphur dioxide and oxygen over heated platinized asbestos in a hard glass tube. Oxygen combines with sulphur dioxide and forms the higher oxide—sulphur trioxide. Platinzed asbestos acts as a catalytic agent. The sulphur trioxide comes out as vapour at the other end of the tube and is collected in a dry flask cooled in a freezing mixture.

Sulphur dioxide+oxygen=Sulphur trioxide.

Note—Platinized asbestos is prepared by dipping a piece of asbestos in a solution of chloroplatinic acid and then heating the mineral in a strong flame when platinum spreads in a fine grey powder on the fibres of the asbestos.

Properties Sulphur trioxide is a colourless, crystalline solid. It has a strong affinity for water and dissolves in water with a hissing noise forming sulphuric acid; much heat is evolved When heated it decomposes into sulphur dioxide and oxygen.

Uses. It is used in the manufacture of sulphuric acid

Sulphuric Acid (Oil of Vitriol)

Preparation. (*i*) By dissolving sulphur trioxide in water. (*ii*) By passing a mixture of sulphur dioxide, air, steam and oxides of nitrogen from nitric acid in

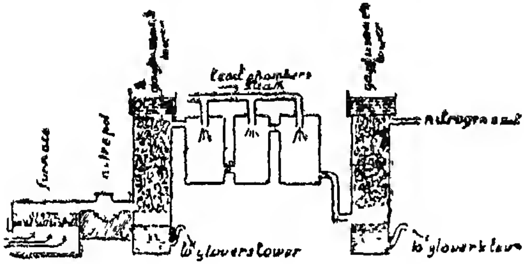


Fig 36

large chambers lined with lead, where sulphur dioxide is converted into sulphur trioxide and is dissolved in water to form sulphuric acid.

Properties. Physical. When pure it is colourless, oily liquid heavier than water, density being 1·84. Commercial acid is pale brown in colour due to the presence of impurities. It boils at 330°C. It turns blue litmus red and has a sour taste. It is soluble in water and much heat is produced during its solution. In order to prepare dilute sulphuric acid pour acid little by little in water and keep stirring the liquid with a glass rod. Never add water to acid as so much heat is evolved that the liquid spurts out and even the vessel may break, producing serious results.

Chemical. Sulphuric acid has a great affinity for water and is, therefore, used as a drying agent for gases like hydrogen, oxygen, air, etc., on which it has no action It is a hygroscopic substance, *i.e.*, takes up moisture from the air when exposed. Sulphuric acid has such a strong affinity for water that it can take up the elements of water (oxygen and hydrogen) from organic compounds, like sugar, wood, paper, etc.

EXPERIMENT 82(*a*) Dip a taper of wood in a tube containing strong sulphuric acid. It is charred.

82(*b*). Prepare a thick syrup of sugar in a dish. Put some strong sulphuric acid in it and stir. Quickly the liquid blackens and a froth is produced. The liquid swells up to fill the disk The black substance left behind is carbon.

Dilute sulphuric acid acts on zinc, iron, aluminium magnesium, forming their respective sulphates and hydrogen. Dilute acid has no action on copper, lead, silver and mercury. Strong sulphuric acid when heated with these four metals or white zinc, iron or aluminium forms their sulphates and sulphur dioxide is given off.

When sulphuric acid is heated with a metallic chloride, it froms metallic sulphates and hydrochloric acid gas,

Sulphuric acid + sodium chloride = Sodium sulphate + hydrochloric acid gas.

When it is heated with nitrates, it produces sulphates and nitric acid ;

Potassium nitrate + Sulphuric acid = Potassium sulphate + nitric acid.

It decomposes carbonates to form sulphates and carbon dioxide :

Sulphuric acid + calcium carbonate = Sodium sulphate + carbon dioxide.

It acts as an oxidizing agent. Sulphur and charcoals

when boiled with it, are oxidized to sulphur dioxide and carbon dioxide respectively ·

Sulphuric acid+carbon=Sulphur dioxide
+carbon dioxide+water.

Sulphur+sulphuric acid=Sulphur dioxide
+water.

Uses of Sulphuric acid. It is the most important acid and is employed in the preparation of hydrochloric acid, nitric acid and in almost all chemical industries. It is used in the manufacture of washing soda, fertilizers, and dye-stuffs, in the separation of gold and silver, in refining petroleum and other oils It is used in storage batteries and fire extinguishers, in explosives and manufacture of indigo.

Test (*i*) Strong sulphuric acid chars wood.

(*ii*) Sulphuric acid or a solution of a sulphate gives a white precipitate insoluble in acids when a few drops of a solution of barium chloride or barium nitrate are added to it ·

Sodium sulphate+barium chloride=Sodium
chloride+barium sulphate (insoluble)

Sulphuric acid+ barium chloride=Barium
sulphate (insoluble) +hydrochloric acid.

Sulphates Sulphates are compounds of sulphuric acid containing a metal, sulphur and oxygen. The important sulphates are :—

(1) Glauber's salt (Sodium sulphate) containing sodium, sulphur, oxygen.

(2) Blue vitriol (Copper sulphate) containing, copper, sulphur, oxygen.

(3) Green vitriol (Iron sulphate) containing iron sulphur, oxygen.

(4) White vitriol (Zinc sulphate) containing zinc, sulphur, oxygen

(5) Epsom salt (Magnesium sulphate) containing magnesium, sulphur, oxygen.

(6) Gypsum (Calcium sulphate) containing calcium, sulphur, oxygen.

The general methods for the preparation of *sulphates* are, by the action of acid—

(*i*) on metals, *e.g.*—

Zinc+dil. sulphuric acid=Zinc sulphate +hydrogen.

Copper+strong sulphuric acid=Copper sulphate+sulphur dioxide.

(*ii*) on oxides of metals, *e.g.*—

Sulphuric acid+calcium oxide=Calcium sulphate+water.

(*iii*) on alkalies, *e.g.*,—

Sulphuric acid+sodium hydroxide=Sodium sulphate+water.

(*iv*) on carbonates, *e.g.*,—

Sodium carbonate+sulphuric acid=Sodium sulphate+carbon dioxide.

(*v*) on chlorides. *e.g.*—

Sodium chloride+sulphuric acid=Sodium sulphate+Hydrochloric acid gas

(*vi*) on nitrates *e.g.*—

Sodium nitrate+sulphuric acid=Sodium sulphate+nitric acid.

All sulphates with the exception of those of barium and lead are soluble in water. Calcium sulphate is only slightly soluble in water causing permanent hardness in water.

(Gypsum—Calcium sulphate). As already stated large deposits of gypsum are found in the Salt Range. Other varieties met with in nature are selenite and alabaster. It has a large amount of water of crystallization which is given off when the mineral is heated. The powder formed is known as Plaster of Paris. This substance sets to a hard mass when made into a thick paste with water. It slightly expands on solidification. It is, therefore, used for making moulds or plaster casts. It is used also in the preparation of black-board chalk. Alabaster is like marble in appearance but much soft and is used in making toys and ornamental vessels.

SUMMARY

Free sulphur occurs in nature in a crude form as brimstone, and in combination it forms sulphides and sulphates

The allotropic modifications of sulphur are (*i*) Rhombic, (*ii*) Prismatic and (*iii*) Plastic Rhombic sulphur is the stable form at the ordinary temperature, the other two convert into it on standing for some time

When any of these varieties is burned, sulphur dioxide is produced

Sulphur dioxide is produced in the laboratory by heating copper with strong sulphuric acid

Properties of sulphur dioxide It is a colourless gas with a characteristic choking smell like that of burning sulphur and a peculiar taste It is heavier than air and soluble in water The solution is known as sulphurous acid whose salts are known as sulphites

It turns blue litmus red, bleaches vegetable colours and is a germicide It is a reducing agent.

Test Paper dipped in acidified potassium dichromate solution is turned green by it Moistened coloured flower is bleached

Uses Used as a bleaching agent, germicide and in the manufacture of sulphuric acid.

Sulphur trioxide When oxygen and sulphur dioxide are passed over heated platinized asbestos sulphur trioxide is produced It is a crystalline substance which dissolves in water to form sulphuric acid

Sulphuric Acid It is prepared by dissolving sulphur trioxide in water When heated, it gives up sulphur dioxide

It is a colourless, oily liquid, has a sour taste and turns blue litmus red, It has a strong affinity for water It chars wood and takes away water from sugar leaving behind carbon It is an oxidising agent It acts upon metals, their oxides, hydroxides, chlorides carbonates and nitrates to form sulphites

Test. Strong acid chars wood

Dilute acid or solution of a sulphate gives a white precipitat with barium chloride or nitrate This precipitate is insoluble in nitric acid and many other industries

QUESTIONS.

1 Where is free sulphur found ? Give the name and composition of five or more native compounds of sulphur.

2. What is (*a*) brimstone, (*b*) rhombic sulphur, (*c*) monoclinic sulphur, (*a*) plastic sulphur ? How is each obtained ?

3. Compare rhombic and prismatic sulphur

4 State the uses of sulphur Describe fully the changes which sulphur undergoes when heated

5 Explain the term allotropy by describing the chief allotropic forms of sulphur. How can you show experimentally that these forms are chemically the same ?

6 Describe the preparation of sulphur dioxide What experiments will you perform to illustrate its chief properties ?

7 Distinguish between the bleaching action of chlorine and that of sulphur dioxide

8 What gaseous compounds of sulphur are you familiar with ? Describe how you would obtain each from ordinary roll sulphur.

9. Describe the chief properties and uses of sulphuric acid

10 What is the action of (*i*) dilute, and (*ii*) concentrated sulphuric acid on zinc, copper, potassium carbonate, nitre, common salt, aluminium and gold ?

11 What is (*a*) gypsum, (*b*) white vitriol, (*c*) een vitriol, (*d*) blue vitriol, (*e*) Glauber's salt, (*f*) oil of vitriol ?

CHAPTER VIII.

CARBON.

The element carbon is very widely distributed in nature. It exists both in the free state and in combination. It occurs in the free state as Diamond and Graphite which are its crystalline froms and as coal which is a non-crystalline form.

In the combined state it is present as carbon dioxide in the atmosphere and in natural waters. It is found in the form of carbonates of metals, chiefly calcium and magnesium, *e g*, limestone, marble, chalk and magnesite. The mineral dolomite contains both calcium and magnesium carbonates. It is present in mineral oils or petoroleum in the form of hydrocarbons. It is a necessary constituent of all organic compounds, *i e.*, compounds met with in the animal and vegetable world.

Organic compounds :—

These compounds contain carbon as an essential part. The number of these compounds is constantly increasing. The branch of chemistry which deals with these compounds is called Organic Chemistry because living plants and animals are largely composed of carbon compounds. It was formerly supposed that these compounds could be formed only by the agency of animal or vegetable life. Quite a number of these compounds have now been prepared by synthetic methods, *e.g.*, Sugar, Indigo,

etc. Some of the important salts are carbohydrates, (*e. g.*, starch, sugar, cotton), alcohol, soap, explosives, dyes, rubber, gum, etc.

The chief points of difference between these and inorganic compounds are :—

Organic.	Inorganic.
1. These contain carbon as an essential constituent.	In these presence of carbon is not necessary.
2. Carbon and hydrogen serve as a starting point for a large number of these compounds.	No single element can serve as a starting point for these compounds.
3 They usually decompose and decay at high temperature.	They do not easily decompose.
4. They are formed as a result of metabolism in living animals.	They are found in the earth's crust as minerals, etc.
5. They are usually soluble in alcohol, ether, ben. zene. etc.	They are generally soluble in water and not in alcohol, etc.
6. They possess the characteristic property of the group.	Every compound has its own properties.
7. Only 20 elements are found in these compounds.	Nearly all the known elements enter in their composition.

The Allotropic Modifications of Carbon. As already stated carbon occurs in three allotropic modifications: (1) Diamond. (2) Graphite and (3) Amorphous Carbon. These various forms though they differ widely in their physical properties are chemically identical. If any variety of carbon is burnt in air or in oxygen, carbon dioxide is formed.

Diamond It is the purest form of carbon. Diamonds are found in various parts of India, in Brazil, South Africa, Australia and United States. When pure, it is colourless, the coloured varieties contain some impurities. The black variety, *carbonado*, is not valued as a gem but is used in glass cutting and drilling instruments. Diamond is the hardest substance known. It has a brilliant lustre on account of its high refractive index and great dispersive power. It is, therefore, used as a gem. It is a bad condutor of heat and elctricity. Its specific gravity is 3 to 3·5.

When diamond is strongly heated without contact with air in an electric arc, it is converted into graphite, but when it is heated in oxygen, carbon dioxide is formed and nothing is left behind. Diamond resists the action of almost all chemical reagents.

Artificial diamonds have been prepared by Moissan, a French chemist. Charcoal is dissolved in molten iron at a very high temperature and the molten iron is suddenly dropped in cold water. The pressure exerted is very great and some carbon crystallizes out as transparent diamond. These diamonds are small in size.

Graphite It is found in large quantities in India Ceylon, Siberia, and California. It is a soft, shiny greyish black substance which is smooth and soapy to the touch. It occurs both in the crystalline and amorphous forms. Its specific gravity is about 2·5. It is a good conductor of heat and electricity. When strongly heated in oxygen, it burns to form carbon dioxide.

Uses. Graphite leaves a black mark when rubbed on paper. So it is used for making lead pencils. In order to use it for this purpose, graphite is purified by grinding and washing to remove the grit. The purified graphite is then mixed with a little washed clay. It is used in the manufacture of plumbago crucibles, electrodes, electric light carbons, etc. It is used as a lubricant for machinery, a coating for iron to prevent rusting.

Note. Graphite is known by the name of black lead or plumbago but it does not contain any lead as was once supposed.

Amorphous Carbon The amorphous varieties of carbon are coal, coke, charcoal, lampblack, animal charcoal, gas carbon. etc.

Charcoal. When vegetable substances like wood are heated in the absence of air, a black substance rich in carbon is left as residue. This is known as charcoal. Wood charcoal is the most important variety and is used in the household as fuel.

Methods of Preparation (a) Cut faggots of wood are piled in a mound in such a way as to leave a

chimney in the centre. (*See* Fig. 37.) The whole is then covered with turf. A lighted faggot is dropped down the chimney to kindle the wood, which burns slowly.

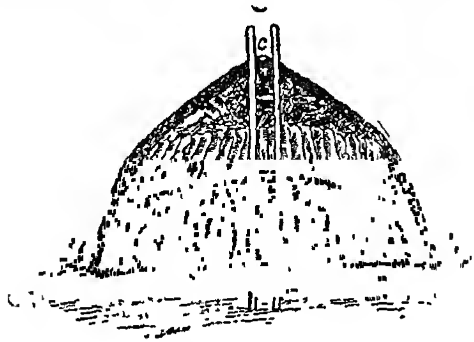


Fig 37

C—Chimney

Holes are left at the bottom to allow air to pass in. When a blue flame is noticed in the chimney after a few days, the holes are closed and the coal is allowed to cool. This process is wasteful as all the volatile products are lost. It has now been replaced by a more economical method described below.

(*b*) Wood is heated in ovens or retorts from the outside, no air being allowed to come in contact with wood. The volatile products are collected and condensed and the inflammable gas is used for heating the retorts. Charcoal is left behind in the retorts. This process is called the *destructive distillation* of wood.

Properties. It is a black, amorphous, porous substance with a specific gravity of about 1·5. On account of its porous nature it floats on water. If the air is removed by an air-pump or by boiling it in water,

then it sinks in water. It is friable. It is a bad conductor of electricity. It has the power of absorbing gases and vapours to a large extent.

EXPERIMENT 83. Invert a cylinder full of ammonia on a trough of mercury. Introduce a piece of recently ignited wood charcoal into the cylinder. Charcoal soon absorbs the gas which is shown by the rise of mercury in the cylinder.

Charcoal absorbs colouring matters from solution.

EXPERIMENT. 84. Slightly colour water with blue litmus. Add powdered charcoal in it, shake and allow it to stand for sometime. Filter the contents. The filtrate will be a clear liquid.

Charcoal under ordinary circumstances is not acted upon by the air but burns to form carbon dioxide when strongly heated It resists the action of moisture and, therefore, the lower ends of poles which are to be stuck in the ground are sometimes charred.

Animal Charcoal or **Boneblack.** It contains only about 10% of carbon and is prepared by the destructive distillation of bones in iron retorts. It is a better absorbent of gases and colouring matters than wood charcoal and is largely used for refining brown sugar.

EXPERIMENT. 85. Boil some solution of brown sugar (*gur*) in a beaker, add powdered animal charcoal and allow it to stand. Filter the the contents. The filtrate will be a colourless liquid.

EXPERIMENT. 86. Boil in another beaker a weak solution of litmus or indigo with animal charcoal and allow it to stand and filter. The filtrate is a colourless liquid.

Lampback is made by burning substances rich in carbon like turpentine, petroleum, etc., in a limited supply of air. The soot is collected by deposition on coarse 'blanket'. It is used for making ink, stove

and shoe polish, etc. It is one of the purest forms of amorphous carbon.

Coal. Mixed with clay and sand coal occurs in nature in large deposits. In India it is found in Dandot mines in the Salt Range, in Bengal and Jharia mines, etc. It is formed by the gradual decay of vegetable matter under the intense heat and pressure of the earth for centuries. There are many varieties of coal between lignite which contains 43%of carbon and anthracite which contains about 87% of carbon. Coal is used as a fuel and in the preparation of coal gas.

Coke Coke is produced when coal is strongly heated in iron retorts from which air has been removed. The volatile products are collected and the greyish black, hard spongy substance—*coke* is left behind as residue. It is used as household fuel as it burns without a flame. It is used as a reducing agent in the extraction of metals from their ores.

The process employed in the preparation of wood charcoal (or animal charcoal or coke) by heating wood etc., in iron retorts in the absence of air is known as *Destructive Distillation*

Gas Carbon It is obtained by the destructive distillation of coal in the manufacture of illuminating gas. It remains in the retort as an extemely hard deposit lining its roof and the sides. It is very pure form of carbon with specific gravity 2·35. It is a good conductor of electricity and is, therefore, used in the manufacture of carbon rods for the arc light.

Chemical properties of Carbon It is inert at ordinary temperature but when strongly heated burns and forms carbon dioxide leaving behind a little ash. It is a reducing agent.

EXPERIMENT. 87. On a piece of charcoal bore a cavity Put in the cavity a mixture of red lead and

powdered charcoal. Heat it with a blow-pipe a bead of lead is formed. Red oxide of lead is reduced into lead :

Lead oxide+carbon=Lead+carbon dioxide.

It unites with sulphur to form carbon disulphide. With metals at high temperature it forms carbides, *e.g.*, calcuim carbide.

SUMMARY

Allotropic forms of carbon (a) Crystalline, (i) diamond, (ii) graphite

(b) Amorphous or non-crystalline, wood charcoal, coal, animal charcoal, lampblack, coke and gas carbon

Diamond is purest carbon. It is the hardest substance known, and has a brilliant lustre It is valued as a gem.

Graphite or blacklead is a soft, shining greyish black substance Good conductor of heat and electricity. It is used for making lead pencil, electrodes, crucibles, and as a lubricant for machines

Wood charcoal is obtained by heating wood out of contact with air in closed retorts or in stacks under earth Animal charcoal is got by distilling bones Wood charcoal is porous and is used for absorbing bad gases Powdered animal charcoal is used for decolorizing solutions wood charcoal is a valuable fuel

Coal is a very impure form of carbon and is used as a fuel It is also used in the manufacture of coal gas which is a mixture of several gases and is used for heating and lighting purposes

Coke is a substance obtained as a residue when coal is distilled in iron retorts In the form of gas carbon it is used for electrical purposes —arc lamps Coke is used as a fuel and in obtaining metals from their ores

Lampblack is carbon deposited by oils, etc , burning in an insufficient supply of air It is pure carbon and is used in black paints and printer's ink.

QUESTIONS.

1 Describe the various allotropic forms of carbon How will you show experimentally that they are really different forms of the same chemical substance ?

2 Campare the properties of diamond with those of graphite How will you show by an experiment that both these substances are carbon ?

3 How is wood charcoal obtained ? State its chief properties and uses.

4 What are artificial diamonds and how are they obtained ?

CHAPTER IX.

OXIDES OF CARBON.

Carbon dioxide:—

Occurrence. This gas occurs in the atmosphere (·03%), its presence there being due to combustion of coal, wood, the breathing of animals and the decay of plants. It is given off in large quantities from the earth in certain places, such as Poison Valley at Java. It is given off in the fermentation of sugar, in the prepartion of wine. It is present in certain natural waters.

Preparation. (1) Carbon dixide is prepared when carbon is burnt in air or oxygen :

Carbon+oxygen=Carbon dioxibe.

(2) The gas is given out and when carbonates like chalk, limestone, marble, etc., are heated strongly as in the case of lime burning :

Calcium carbonate→Calcium oxide+carbon dioxide
(Limestone) (Lime)

This method is employed for the manufacture of gas on a large scale.

(3) Pure carbon dioxide is prepared by heating pure sodium bicarqonate :

Sodium bicarbonate=Sodium carbonate+ water+carbon dioxide.

(4) On a large scale for manufacturing purposes carobn dioxide is obtained as a by-product from fermentation, during the manufacture of alchol. The gas is collected in wrought iron cylinders in which it is liquefied under pressure and is sold in the liquid form.

(5) *Laboratory Method.* In the laboratory is most easily prepared by the action of an acid such as hydrochloric acid upon a carbonate such as limestone or marble:

Calcium carbonate + hydrochloric acid
(Marble)
= Calcium chloride + carbon dioxide + water.

EXPERIMENT 88. Fit up the apparatus as shown in the Fig. 38; see that the apparatus is air-tight. Put few pieces of white marble into the flask, cover these with water and put in the cork See that the lower end of the thistle funnel dips under water. Pour strong hydrochloric acid down the thistle

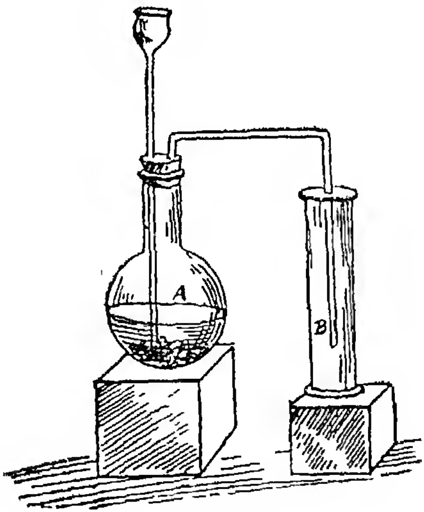


Fig 38

A—flask containing marble pieces.
B—carbon dioxide and dilute hydrochloric acid.

funnel. A brisk effervescence takes place. Collect the gas by the downward displacement of air. In order to find out whether the cylinder is full, take a lighted match stick near the mouth of the jar. If it goes out, the jar is full. Collect a few cylinders of the gas and perform the following experiments with them.

EXPERIMENT 89. Examine a jar full of carbon dioxide. See that it has no colour; smell it—it has a faint pungent smell; taste it—it has got an acid taste A bottle of soda water which is a solution of the gas in water shows the taste of the gas.

EXPERIMENT 90. Invert a cylinder full of the gas in a trough of water and allow it to stand there for some time. Notice the small rise of water in the cylinder The gas is soluble in water.

EXPERIMENT 91. Invert a jar full of gass over a empty jar as shown in the diagram. Cover the lower cylinder. Introduce a burning match stick in each of the cylinders. It is extinguished in the lower cylinder which was formerly empty but continues burning in the cylinder which was previously full of the gas. This shows that the gas has gone from the upper to the lower cylinder. Carbon dioxide is heavier than air and can be poured like water from one vessel to another.

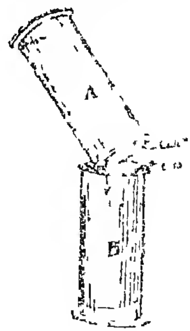


Fig 39

EXPERIMENT 92: Introduce a lighted candle in a jar. It at once goes out. Pour the gas as shown in the diagram over the flame of a lighted candle. The candle at once goes out.

Fig 40

EXPERIMENT 93. Put a moistened blue litmus paper in a jar containing the gas; it is changed into red colour.

EXPERIMENT 94 (*a*). Put some lime water in a jar and shake The lime water becomes milky.

(*b*). Pass the gas from the delivery tube through some lime water. Observe that it becomes milky. When the gas is continued to pass, the milkiness disappears. Disconnect the appartus. Boil the clear liquid. Milkiness again appears. Put a few drops of hydrochloric acid in this milky liquid. Note the effervescence and disappearance of milkiness. Soda water may be used for this experiment intead of the gas.

Properties. Physical. Carbon dioxide is a colourless gas with a faint pungent smell and acid taste. it is about $1\frac{1}{2}$ times as heavy as air. It is soluble in water. Soda water is a solution of this gas under pressure in water. The gas can be liquefied and solidified into a snow-like solid mass by lowering the temperature and increasing the pressure.

Chemical. It does not burn nor supports combustion. Air which contains $2\frac{1}{2}\%$ by volume of this gas does not allow substances to burn in it though there

is about $18\frac{1}{2}\%$ of oxygen still present in air. It does not support respiration. It is not poisonous and animals die in it from suffocation. It is slightly acidic in properties. Its solution in water is a weak acid—Carbonic acid :

Carbon dioxide+water=Carbonic acid.

This is why it is sometimes called carbonic acid gas. It is inert but some substances like magnesium and potassium burn in it decomposing carbon dioxide into carbon and oxygen. They combine with oxygen and carbon is deposited in black particles on the sides of the vessel :

Magnesium+carbon dioxide=Magnesium oxide+carbon

Potassium+carbon dioxide=Magnesium carbonate+carbon.

When passed through lime water, it forms calcium carbonate white which being insoluble in water renders it milky :

Calcium hydroxide+carbon dioxide=Calcium carbonate+water

When more carbon dioxide is passed the lime water becomes clear on account of the formation of calcium bicarbonate which is soluble :

Calcium carbonate+carbon dioxide+water =Calcium bicarbonate+water.

When the solution is boiled, the water again becomes milky on account of the escape of carbon dioxide and formation of insoluble calcium carbonate from calcium bicarbonate :

Calcium bicarbonate $\xrightarrow{\text{heat}}$ Calcium carbonate+carbon dioxide+water.

When we pour hydrochloric acid now in this milky liquid, carbon dioxide is produced and the water becomes clear.

(i) Calcicum carbonate+hydrochloric acid

=Calcium chloride +carbonic acid.

(ii) Carbonic acid being unstable at once gives up carbondioxide :

Carbon acid→Carbon dioxide+water.

It is decomposed by the green cololuring matter of plants in the presence of sunlight, carbon being assimilated and oxygen set free.

Test. It turns lime water milky.

Uses. (1) It is used as a fire extinguisher. Fire extinguishers generally seen hanging in big buildings contain sodium carbonate and a tube containing hydrochloric acid When the two are mixed by breaking the tube, the gas is evolved and is used in putting out the fire.

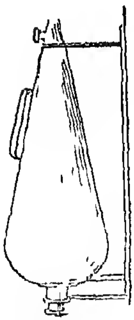


Fig. 41

(2) It is used in ærated waters—Soda water.

(3) Solid carbon dioxide (carbonic acid snow) is used as medicine—a cure for local sore and in freezing mixtures.

(4) It is used in the manufacture of white lead and sodium carbonate (washing soda).

Carbonates Carbonates are the salts of carbonic acid. The most important carbonates are those of calcium and magnesium. They are very common in nature as limestone, chalk and marble, and magnesium carbonate in the form of dolomite Shell, corals, pearls, etc., are also calcium carbonate.

There are two crystalline varieties of calcium carbonate : (i) Calcite or Iceland Spar—it is a doubly refracting medium and (ii) Aragonite.

Sodium carbonate or washing soda is another important carbonate.

Manufacture of Quicklime (chuna). The limestone (*pathar chuna*) is packed in a lime kiln lined with brick work. (*See* Fig. 42.) Limestone and coal or wood are placed in the kiln in alternate layers. The fuel is burnt in a free supply of air entering through a large opening at the base. Carbon dioxide mixed with air escapes, and lime is left behind. This lime is, however, mixed with fuel ashes and hence impure.

In some cases the limestone is arranged in an arch over the fire below, the fuel being kept separate from the stove. The burning goes on for about two days, after which the kiln is allowed to cool and the lumps of quicklime are removed In this process there is much waste of fuel but the lime obtained is pure.

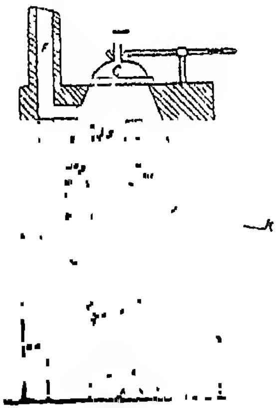


Fig 42

K—lime kiln containing coal and limestone.
C—Iron grid through which lime falls down intoP, pit
F—Chimney.

The lime thus obtained is called quicklime. It is a white amorphous substance which melts at a very high temperature When heated in oxyhydrogen flame it emits a bright light known as limelight. It

melts only in the electric furnace. When a small quantity of water is poured on lime, great heat is given off, and water is at once converted into steam, the stone swells up and crumbles to a soft, dry powder—slaked lime. The process is known as the 'slaking of lime'. The lime chemically combines with water to form calcium hydroxide or slaked lime :

Calcium oxide + water = Calcium hydroxide.

Quicklime is used as a dehydrating agent for absorbs moisture and it also absorbs carbon dioxid A dehydrating agent removes moisture from other sub stances It is used in purifying the water of wells an as a disinfectant.

Slaked lime is a white amorphous powder, sparin soluble in water It is less soluble in hot water than cold one The solution in water is known as li water. It is slightly alkaline and absorbs carb dioxide forming calcium carbonate—chalk Slak lime when mixed with water (not sufficient to disso it) forms a white paste—milk of lime. This wh exposed to the atmosphere for a few days sets to hard substance, hence lime is used in the manufac of mortar. It is used for white-washing purposes, purifying sugar and removing hair from hi It is used in the manufacture of washing soda bleaching powder It is used as fertilizer.

Mortar. Mortar is made by mixing slaked with three or four times its bulk of sand and making whole into a paste with water. The moisture is porated or absorbed by the bricks and the m becomes hard Hardening is further helped by atmospheric carbon dioxide which which slowly conve

into calcium carbonate. There is a slight combination between lime and sand. Sand also makes the mass porous so that carbon dioxide can penetrate below the surface.

Cement. It is made by burning a mixture of clay, limestone and gypsum. Clay is a mixture of aluminium oxide and sand (silica). On mixing with water, cement sets to a hard mass owing to the formation of silicates. There are many kinds of cements in the market. Portland cement, Banil cement, Roman cement, etc. Cement is now manufactured on a large scale in India.

Concrete is a mixture of cement and gravel. Houses and bridges are often built over a framework of iron rods. It is then called reinforced concrete.

Stalactites and **Stalagmites.** As already stated, water containing carbon dioxide dissolved in it can hold calcium carbonate in solution. Such water when it falls in drops from the roofs of caves in limestone districts loses the carbon dioxide by evaporation and deposits this insoluble calcium carbonate in the form of pendants, hanging from the roof, formed of several concentric layers. They are known as stalactites

The drops falling on the floor of the cave also deposit the insoluble carbonate and pillars grow up from the floor upwards and may meet the stalactites. They are known as stalagmites.

Stalactites and stalagmites are not necessarily of calcium carbonate though best specimens are of this substance. These formations of salt are met in the salt mines.

Carbon Monoxide.

Occurrence It occurs in the coal gas and is seen burning with a blue flame over a clear fire of glowing charcoal The smoke given out by motor engine also contains this gas

Preparation (1) It is prepared by the reduction of carbon dioxide by charcoal .

carbon dioxide=carbon+carbon monoxide

Charcoal is packed in a hard glass tube open at both ends and strongly heated Carbon dioxide from the gas apparatus is passed over the heated charcoal Some of the gas is reduced and a mixture of carbon dioxide and carbon monoxide escapes from the other end Carbon monoxide is collected over a solution of caustic potash (*See* Fig 43) which absorbs carbon dioxide, but has no effect on carbon monoxide

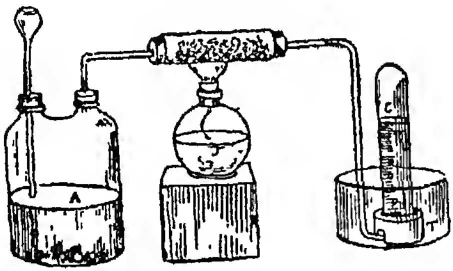


Fig 43

A—carbon dioxide apparatus

C—carbon monoxide

T—trough containing caustic potash solution

(2) In the laboratory it is prepared by heating sodium formate with concentrated sulphuric acid Sulphuric acid removes the

elements of water from the formic acid and leaves behind pure carbon monoxide, which is collected over water (*See* Fig 44).

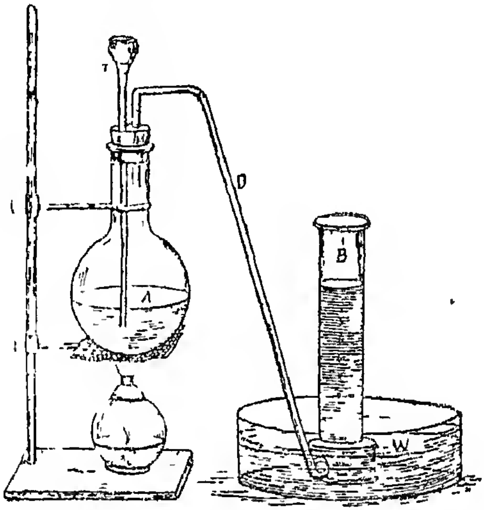


Fig 44

A—flask containing sodium formate and strong sulphuric acid

B—carbon monoxide

(3) Oxalic acid crystals when warmed with sulphuric acid give out equal quantities of carbon dioxide and carbon monoxide The gases are separated by passing the mixture through a solution of caustic soda;

Caustic soda + carbon dioxide = Sodium carbonate

Properties—Physical. It is a colourless gas with a peculiar faint smell It is very poisonous ,hence great care must be exercised in performing experiments with this gas or in its preparation It is little lighter than air and very lightly soluble in water

Chemical.—It is a combustible gas, burns with a blue flame in air or in oxygen. When it burns it forms carbon dioxide. It does not support combustion. It is a reducing agent and is used in the extraction of iron from its ores,

Iron oxide + carbon monoxide = Carbon dioxide + iron

It is neutral to litmus and has no effect on lime water.

Note. Rooms should never be closed with charcoal or coal burning inside. The coal should be well ignited outside and then brought into the room, otherwise carbon monoxide will render the air foul. Many people have lost thier lives by sleeping in closed rooms with charcoal glowing in.

Coal gas.—

Preparation. EXPERIMENT 95. Set up the apparatus as shown in the diagram and see that the

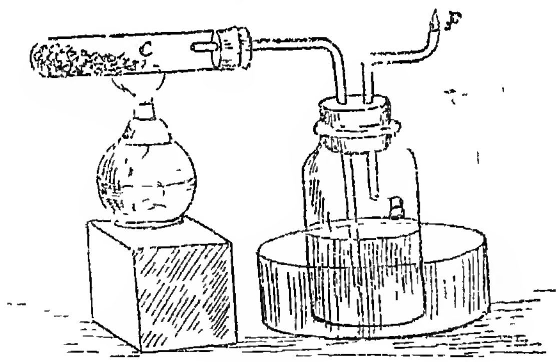


Fig 45

C—powder coal.
B—vessel containing water.
F—gas burning.

connections are air-tight. Wrap wire gauze round the hard glass test tube and fill two-thirds of it with powdered coal. Place the jar in cold water. Heat the test tube from end to end. In a short time a gas will be seen bubbling through water in the jar and coming out of the jet. In the jar there will be a deposit of tar and ammonia will dissolve in water The gas which comes out of the jet burns when burning taper is applied to it. The gas which comes out is known as coal gas. It can be collected over water like hydrogen or oxygen. Coke is left in the test tube.

Properties. It is a colourless gas with a smell like that of burning coal. It is lighter than air and sparingly soluble in water. It is not a single gas but is a mixture of hydrogen, methane, carbon dioxide, acetylene, benzene, nitrogen, carbon dioxide and oxygen. The last three gases are the impurities.

Uses. It is used for heating and lighting purposes and in filling balloons.

CHAPTER X.

BURNING AND BREATHING IN AIR.

We have already seen that when substances burn in air, they combine with oxygen of the air. The burning of wood, candle, charcoal, etc., are familiar examples of burning in everyday life.

Substances like these when heated give off smoke, being changed first into a black residue. When the heating is continued in the presence of air, the black residue disappears leaving behind a small amount of ash.

The smoke and the vapour given off during burning tell us much about the composition of the substances that burn. As soon as wood or other substances containing carbon are heated in a test tube, moisture is deposited on the cool parts of the tube. When wood is burnt in contact with air, carbon dioxide is produced.

Burning of a candle.—

EXPERIMENT 96*a*. Over a burning candle invert a cold jar which has been dried carefully. Notice that the inside of the jar becomes dim with moisture

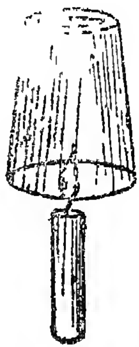


Fig. 46.

EXPERIMENT 96*b*. In a thin glass beaker put some ice cold water. Dry the outer side and bottom of the beaker. Keep a burning candle under the bottom of the beaker for a short time. The bottom becomes covered with droplets of water.

EXPERIMENT 97. Fix a candle in a deflagrating spoon and having lighted it, lower it in a jar. Observe that the flame of the candle becomes dim and soon goes out. Water collects on the inside of the jar. Take out the candle and cover the jar with a glass disc. Quickly lower into the jar a burning taper, it at once goes out. Pour in a little lime water and shake it up in the jar; notice that it is turned milky.

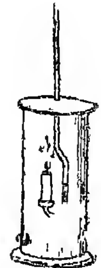


Fig. 47

The above experiments thus clearly show that when a candle is burned in air, water vapour and carbon dioxide are produced.

In fact this is true in all cases of burning organic substances (i.e., containing carbon) in air, viz., water vapour and carbon dioxide are produced.

Breathing (respiration) of animals in air It has been learnt that oxygen is essential for animal life; without it animals would die. In the process of breathing the animals take in air through the nose and mouth to the lungs. In the lungs the oxygen of the air passes into the blood and is carried to the different parts of the body. It is slowly used up in burning up the waste tissue of the body. These tissues contain, besides other elements, carbon and hydrogen which are oxidized into carbon dioxide and water respectively. The heat produced during the oxidation of the tissues supplies the necessary heat and energy to the body. The following experiments clearly illustrate the above facts.

Experiment 98. Blow through a glass tube into some clear, freshly made lime-water in a beaker. Notice it is rurned milky.

Experiment 99. Blow on a dry, clean surface of a looking glass. It becomes misty because the moisture present in the expired air deposits on it.

The warmth of the breath may be left by blowing on the hands.

The above experiments clearly show that both in burning and breathing the same substances—carbon dioxide and water vapour—are produced. In both cases, heat is produced as a result of the chemical changes brought about. In the cases of burning, however, the heat produced is great and produces a flame.

How green plants use the air. An animal in breathing uses oxygen from the air and gives back to the air carbon dioxide as a waste product. The air out-of-doors, however, never contains much of carbon dioxide in it because plants remove this gas from the air under certain circumstances.

Experiment 100. (*a*) Take a bunch of water weeds, and put it in a jar nearly filled with water having carbon dioxide dissolved in it. Soda water which has been left uncovered for some time may be used. Cover the weeds with a glass funnel and invert over the funnel a test tube full of the same water as shown in diagram. See that there are no bubbles of air in the test tube or sticking to the weeds. Place the apparatus in bright sunlight for an hour or two.

Bubbles of a gas collect in the tube and are seen sticking to the leaves. Remove the test tube and test the

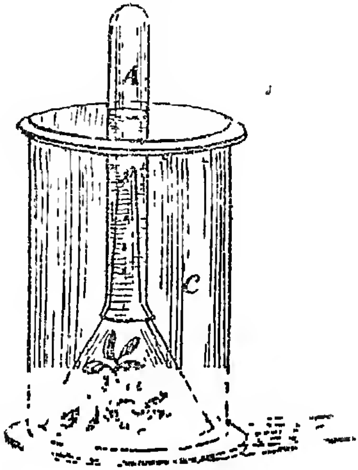


Fig. 48.

C—water containing carbon dioxide.
A—oxygen

gas by a glowing splinter. It is rekindled. This shows that the gas is oxygen.

(*b*) Repeat the same experiment but keep the apparatus in some dark place. No bubbles will be formed in this case.

It is clear from these experiments that green plants in the presence of sunlight decompose carbon dioxide into its parts. They assimilate carbon which is necessary for their growth and set free oxygen. It is

in this way that the amount of carbon dioxide in the atmosphere is kept constant.

The decomposition of carbon dioxide and liberation of oxygen by green vegetables in the presence of sunlight is known as assimilation. This is quite different from respiration. In plants both these processes are going on side by side in the day time but the amount of carbon dioxide given out by respiration is very little as compared to the liberation of oxygen by assimilation. At night, however, the process of assimilation stops but respiration goes on. It is for this reason why it is not advisable to keep plants at night in a sleeping room. The plants instead of purifying the air as is the case in day time, render it impure by breathing out carbon dioxide.

Breathing at high altitudes. You have already learnt that the air pressure decreases with increase in height above sea-level. Hence it is clear that the presence of oxygen in the lungs and in the blood must also decrease, and, if the decrease is considerable, suffocation results. It is possibe to become acclimatized to high altitudes, as, for example, in Tibet which is about 15,000 ft. about sea-level Generally rapid ascent results in discomfort, if not worse. To overcome the defect due to the lack of oxygen at great heights, an artificial supply of the gas is often used by mountain climbers.

SUMMARY IX & X

Oxides of carbon are, (i) carbon dioxide, (ii) Carbon monoxide

Carbon dioxide is obtained by the action of dilute hydrochloric acid on marble pieces (calcium carbon). The gas is collected by the upward displacement of air

Properties. Colourless, heavier than air, has a faint pungent smell and acid taste. Soluble in water, the solution is called

carbonic acid—a weak acid, it does not burn nor help burning. Animals die in it from suffocation, *i e*, want of oxygen

It is inert Magnesium and potaium burn in carbon dioxide decomposing it into oxygen and carbon Turns lime water milky Green plants in presence of sunlight locompose carbon dioxide, carbon being assimilated and oxygen set free.

Uses. Used as a fire extinguisher, in the manufacture of aerated water, washing soda and white lead, as a cure for local sores (Carbonic acid snow)

Carbon monoxide obtained by reduction of carbon dioxide by charcoal or by the action of strong sulphuric acid on sodium formate

Properties No colour, and is very poisonous Has no action on lime water and litmus It burns but does not help burning

Coal gas is a mixture of several gases and is used for heating and lighting purposes

When organic substances (candle) burn in air, water vapour and carbon dioxide are produced

In respiration animals take in oxygen of the air give out carbon dioxide, green plants in the presence of sunlight decompose carbon dioxide into carbon and oxygen which is set free

A carbonate is a substance which gives carbon dioxide when heated or acted upon by an acid

The chief carbonates found in nature are chalk, limestone, marble shells, corals, pearls, etc. Calcite and washing soda are crystalline forms of calcium carbonate and sodium carbonate respectively

Lime is obtained by heating limestone in lime kilns Freshly burnt lime is called quicklime Slaking of lime is a chemical change It is used as a dehydrating agent and for producing lime light Slaked lime is slightly soluble in water forming lime water It is used in the manufacture of mortar, washing soda, bleaching powder, and for white washing, etc

QUESTIONS

1 Write a brief description of the different forms of of carbon By what experiments would you prove that they are relly different forms of the same element ?

2 Compare the properties of diamond with those of graphite How will you show that both these substances are carbon ?

3 Explain the terms 'crystalline, and 'amorphous' using carbon as an example

4 What is amorphous carbon ? Write what you know about the different important properties of amorphous carbon

5 How is coal formed in nature ? Give its chief properties and uses

6 What is meant by destructive distillation ? What substances can be obtained by the destructive distillation of coal ?

7 Give an account of the allotropy of carbon How are artificial diamonds made ?

8 How is charcoal prepared on a large scale from wood ? What are its properties and uses ?

9 Describe the prepartion of carbon dioxide, and give experiments to illustrate its properties

10 A glass jar is supposed to contain either oxygen, nitrogen, hydrogen or carbonic acid gas How will you ascertain its nature ?

11 A candle is burnt under a gas jar and eventually the flame goes out How would you show the presence of carbon dioxide and water vapour in the jar ?

Whence do you suppose that different elements forming carbon dioxide and water come ? Give reasons for your answer

12 How is carbon monoxide prepared ? How does it differ from carbon dioxide ? How can one be changed into the other ?

13 Carbon dioxide does not support burning but potassium and magnesium burn in it Why is it so ? Describe the action which takes place in each case

14 What chemical changes take place when carbon dioxide is passed through lime water (*a*) in a small quantity, (*b*) in excess ? What happens when the solution in (*b*) is boiled ?

15 You are given some lime water and asked to blow into it Describe the changes you observe, and name the substances produced

(*P U*,, *Matric*, 1920)

16 How can you prove the presence of carbon dioxide in air ? Describe an experiment to show the importance of carbon dioxide to plants (*P U*,, *Matric*, 1919)

17 Describe fully the manner which the breathing of animals and plants effect composition of the air

18, What is a carbonate ? Name the chief varieties of calcium carbonate and give the properties and uses of each

19 What elements are contained in chalk ? What happens when it is strongly heated ?

How is lime manufactured on an a large sale ?

23 What is (*a*) lime , (*b*) slaked lime , (*c*) lime water, (*d*) mortar, (*e*) cement ? How is each obtained ?

21 Describe the preparation and uses of coal gas

22 Describe experiments to show that burning and breathing are similar processes In what respects do they differ from each other ?

23 Give experiments to shew that —

1 Sugar, potato, etc contain carbon.
2 Carbon dioxide is denser than air
3 Carbon doixide is produced when lime juice is poured over chalk
4 Soda water cantains carbon dioxide

24 How are stalactites and stalagmites formed in nature ?

COMBUSTION AND FLAME

Combustion. Combination of substances with oxygen is called oxidation. It is a slow process and only a small amount of heat is produced. Combustion is a form of oxidation attended with heat and light In general the term combustion is not only applied to burning of substances in air but is also used to describe any chemical reaction that takes place with sufficient energy to produce heat and light. Any substance that burns in air is called *combustible*, and air and oxygen are called *supporters of combustion.* The burning of coal gas in the air is in reality, simply, a chemical reaction taking place between the carbon and hydrogen of the gas and the oxygen of the air, just as the combustion of copper and sulphur is a vigorous chemical action between the two elements.

The terms, *combustible* and *supporter of combustion,* are interchangeable, for by bringing suitable changes in the experiments, the combustible body may be made the suppporter of combustion and *vice versa.* Air burns in an atmosphere of coal gas and *vice versa.*

EXPERIMENT 101. Fit a wide tube with a cork. (Fig 50.) On top of it place a piece of asbestos having a small hole in it. Pass coal gas into the apparatus and burn it as it comes aut from the hole. Push up the tube through which the air enters till it reaches the coal gas flame. Withdraw it and

the air will continue to burn in an atmosphere of coal gas. At the top coal gas is burning in air.

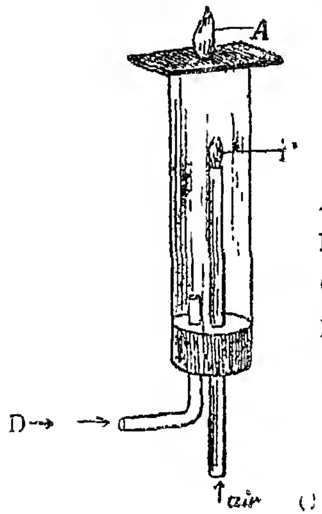


Fig. 50.

A—coal gas burning in air.

B—air burning in coal gas.

C—air enters.

D—coal gas enters.

From the above experiment it is evident that the terms combustible and supporter of combustion are interchangeable. This process is sometimes called *inverted combustion*, but the chemical changes that take place between the two substances are the same in either case. In the same way oxygen will burn in an atmosphere of hydrogen.

EXPERIMENT 102. Invert a large gas jar and fill it with hydrogen. Apply a light to the mouth of the jar. The hydrogen will burn.

Immediately pass up into the jar a tube from which oxygen is being passed. The gas burns as it passes through the hydrogen flame and continues to burn in the atmosphere of hydrogen.

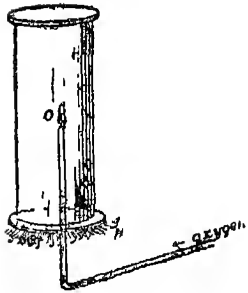


Fig 51

H—hydrogen.

O—oxygen burns.

Flame. When chemical combination takes place between gases or substances which give off inflammable vapours on heating, such as wood or coal, flame is produced. If one or both of the substances taking part in combustion are solids, no flame is produced Magnesium burns in oxygen with a dazzling light. It becomes extremely hot and begins to glow but there is no flame Iron becomes white hot when burnt in oxygen and again there is no flame In these cases the solids are not vaporized and hence no flame. But when a candle is lighted the wax is melted and then changed into a vapour. In this state it unites with the oxygen of the air forming a flame which in reality is the space in which the chemical change is taking place. Similarly sulphur, phosphorus, spirit, kerosene oil, etc., are first changed into vapours, before they actually begin to burn, and produce flame.

The structure of a candle flame Light a candle and examine its flame carefully The following parts can be seen in the candle flame :—

1. *The black inner zone* surrounding the wick. It consists of unburnt gas, or vapour of partly

decomposed wax No combustion takes place in this zone of the flame.

EXPERIMENT 103. Hold one end of a small narrow glass tube of about 5mm bore in the candle flame at this place. Notice the unburnt gas passes up the tube and comes out at the other end. Apply light to this end, the gas burns

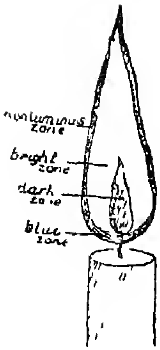


Fig 52.

EXPERIMENT 104. Lower a piece of stout white paper into the flame and quickly remove it. The paper is charred in the form of a ring inside of which is unchanged.

2. *The luminous zone.* Outside the dark inner zone is found the brightest part of the flame. This is the largest part of the flame and is the zone from which soot is deposited on a cold object. In this zone the gases are partly burnt, and the unburnt solid particles of carbon from the wax are heated strongly and make the flame bright.

EXPERIMENT 105. Hold the end of a glass rod in the bright part of the flame Notice black soot deposits on it

3. *Non-luminous zone.* Outside the luminous zone is an almost colourless zone. In this part burning is completed. It is the hottest part of the flame.

EXPERIMENT 105–A. Dip one end of a moistened glass rod in the powder of copper sulphate and hold it just touching the bright zone. A green mantle surrounding the bright zone is seen.

The non–luminous zone becomes luminous by introducing solid particles of copper sulphate or common salt in it.

4. *The non-luminous blue zone* at the base of the flame. Here the combustion is complete due to the presence of plenty of oxygen of the air

The flame of an oil lamp or an ordinary bright gas flame if examined, is found to consist of the same four parts as a candle flame although the shape may be different.

The Bunsen burner is the simplest type of a gas burner. Coal gas or oil gas can be burnt in it with nearly complete combustion. The coal gas supplied through a side tube comes out through a small opening O and at once mixes with the air, which enters through hole at H. The mixture of gas and air passes up the tube and being ignited burns with a non-luminous flame at the top. If we shut off the air supply at the botton by closing the holes the flame becomes luminous.

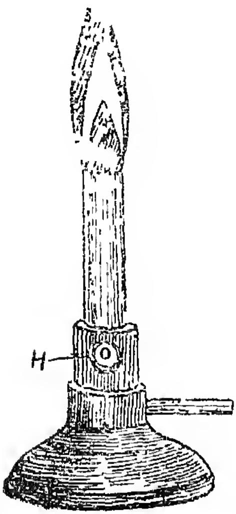


Fig 53

When the air holes of the Bunsen burner are open the flame is made up of two parts; (i) a zone of unburnt gases at the base—bluish in colour, (ii) a faint bule cone surrounded by an almost invisible region of complete combustion. The air entering at the base causes combustion to go on inside the flame as well as outside. This causes the carbon to be burnt away completely to carbon dioxide without depositing any soot.

Luminosity of flames. We have already seen that the bright zone of a candle flame contains unburnt particles of carbon which deposit in the form of soot, and that the Bunsen flame deposits no soot. Thus it may be said that a flame that deposits soot is luminous and a flame that deposits no soot is non-luminous. Hence luminosity of a flame is generally due to the presence in it of small particles of a solid which become sufficiently hot to emit light.

If a piece of lime is held in a bunsen flame, it becomes white hot, and a very strong light known as lime-light is produced. This principle, that solid substances placed in a non-luminous flame cause it to give out light, is used in the ordinary incandescent burner. In the Kitson lamp, for example, intensely bright flame is produced due to the heating of incandescent mantle which is coated with fine partlcies of the oxides of two rare metals, cerium and thorium.

Ignition temperature. A combustible substance will not burn unless heated to a certain temperature. This temperature is called its ignition temperature. Thus the ignition temperature of yellow phosphorus is about 34°C.

If by some means the ignition temperature of a burning substance is lowered, it stops burning.

EXPERIMENT 106. Make a coil of a stout copper wire. Lower it gently into a candle flame. The flame goes out. The copper coil should be close and sufficiently long.

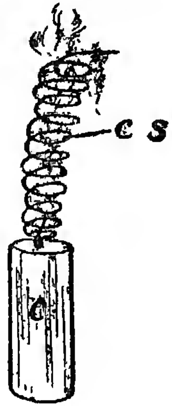


Fig. 54.
C.S—copper coil
C candle.

The heat of the flame is taken away by the copper-wire so rapidly that the temperature of the burning gas in the flame is lowered belowed the ignition temperature and hence the flame goes out.

SUMMARY

Combustion is any chemical reaction that takes place with suffcient enegry to produce heat and light Oxidation is a slow process, while combustion is very rapid

Flame may be defined as a space in which chemical reaction between gases takes place, accompanied by the prodution of heat and light If one or both of the substances taking part in combustion are solids, no flame is produced

A candle flame is made up of four parts (i) A dark inner zone of no combustion round the wick. (ii) A bright zone of incomplete combustion (iii) A non luminous zone of complete combustion (iv) blue zone at the base—complete compustion

Some flames are luminous, some are non-luminous The luminosity of a flame is due to the presence in it of small particles of solid which become sufficiently hot to emit light

Ignition temperature of a substance is the temperature to which it must be heated in order to make it catch fire

QUESTION

1 Write what you know about oxidation and combustion, Is combustion oxidation ?

Explain clearly the terms "combustible" and "supporter of combustion "

2 What is an inverted combustion ? Describe experiments to illustrate your answer,

3 Describe the structure of a candle flame What experiments will you perform to show the presence of unburnt vapour in the inner dark zone of the candle ?

4 Describe the structure of a candle flame and give a brief account of the experiments that you would perform to show the change which takes place when a candle burns in air (*P U* ,*M &S L C* ,1925,)

5 Describe the *construction of a Bunsen burner* When is the bunsen flame (*i*) luminous, (*ii*) non luminous, and why ?

6 What is the luminosity of a flame due to ? Explain Ignition Temperature

CHAPTER XI.

PHOSHPHORUS AND MATCHES.

Occurrence. It has been already stated in a previous chapter that phosphorus has got a strong affinity for oxygen and that it begins to burn when exposed to air. You have seen that phosphorus is kept under water to protect it from air. It is, therefore clear that phosphorus cannot occur free in nature. It occurs always in the combind state, chiefly as phosphates. Calcium Phosphate is a chief constituent of animal bones and teeth. Complex compounds of phosphorus are present in muscles, nerves and brain of animals and are found also in plants. Amongst foods, egg yolk, beans, nuts, peas and wheat are rich in phosphorus. Calcium phosphate is present in all fertile soils.

Extraction 1. Formerly phosphorus was obtained from bone ash—the residue left after the destructive distillation of bones. This consists chiefly of calcium phosphate. Bone ash was treated with sulphuric acid in large wooden vessels to get phosphoric acid :

Calcium phophate+sulphuric acid=calcium sulphate+phosphoric acid.

The phosphoric acid thus formed was separated by filtration, and was concentrated to a syrup. The concentrated liquid was then mixed with either sawdust or powdered charcoal or coke and heated in iron vessels to dry the mass. The dried mass was further heated in clay retorts to a white heat. Phosphorus distilled over and was condensed under water. It was further purified and cast into sticks.

Now-a-days when electric power is cheap, phosphorus is prepared by heating strongly a mixture of bone ash or other mineral phosphates, coke and sand in an electric furnace. Phosphorous vapours which escape near the top are condensed by cooling and the solid is further purified by melting under a solution of chromic acid and cast into sticks of white phosphorus, the common form.

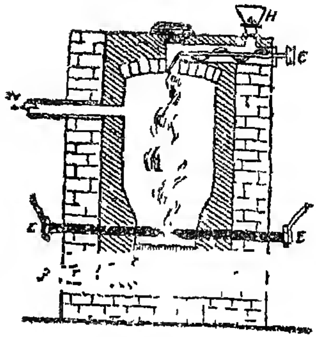


Fig 55.

F—furnace, E—electrodes. P.V.—phosphorus vapours. C—conveyer. H hopper.

Properties. Phosphorus exists in two allotropic forms .—

1. White or yellow phosphorus —the common form

2. Red phosphorus—the stable form at the ordinary temperature.

White Phosphorus. It is almost colourless when pure but turns yellow on exposure to light, hence called yellow phosphorus. It is a translucent solid, like wax in appearance. It is soft and can be easily cut with a knife. It is kept under water (should be cut under water, because it readily combines with oxygen. It melts at 44°C. and boils at 287°C. Its specific gravity is 1·83 to 1·85. It is only very slightly soluble in water , is dissolve readily in ether, olive oil, benzene and is very soluble in carbon disulphide. If its solution in carbon disulphide is poured over a filter paper and exposed to air, it catches fire It should not be touched by hand because it produces burns which are difficult to cure,

EXPERIMENT 107. Cut a small piece of phosphorus under water and put it in a dish containing carbon disulphide. Place a filter paper on a tripod and pour the solution on the filter paper. Below the tripod place a large dish to catch the drops of the solution. Carbon disulphide soon evaporates and phosphorus is left behind on the filter paper in a finely divided state. Soon the filter paper catches fire.

Both yellow phosphorus and its vapour are highly poisonous. Phosphorus readily undergoes oxidation and forms white fumes of phosphorus pentoxide when exposed to air. If kept in the dark, the oxidation is accompanied by a faint green glow, known as phosphorescence. It burns at 34°C, giving dense white fumes of phosphorus pentoxide.

Red Phosphorus. It is prepared by heating fused white phosphorus in a large cast iron retort provided with a cover. The temperature is kept at 230°C. The yellow phosphorus is converted into red phosphorus.

It is a violet red coloured powder. It was formerly regarded as amorphous but is really crystalline. It is insoluble in carbon disulphide. Its density is 2.1. It does not oxidize when exposed to air nor does it glow in the dark. It is not poisonous. It burns at about 240°C, forming white fumes of phosphorous pentoxide. It is not so active as white phosphorus. It is the stable form at the ordinary temperature, for white phosphorus passes slowly into red phosphorus, even at ordinary temperature, when exposed to light.

Both are, in reality, the different forms of the same element, as is proved by the fact that when equal quantities of both are burned, the same amount of phosphorus pentoxide is formed in each case.

Uses. Both the varieties are used in the manufacture of matches. The use of the yellow variety is not allowed by law because it is more poisonous and injurious to factory labourers. Yellow phosphorus is ued as a rat poison.

Oxides of Phosphorus. There are two important oxides of phosphorus, *viz.*, Phosphorous trioxide and Phosphorous pentoxide

Phosphorus trioxide is formed by burning phosphorus in a limited supply of air. It is formed mixed with phospor us pentoxide. It is a snow-white waxy looking crystalline solid. It has an unpleasant garlic like smell. It is very poisonous. It slowly dissolves in cold water and forms phosphorous acid. The salts of phosphorus acid are known as phosphite, *e.g.*, when this acid acts upon caustic soda, it forms sodium phosphite ;

Phosphorus acid+sodium hydrate
(caustic soda)

=Sodium phosphite+water.

When exposed to air or oxygen, it oxidizes into phosphorus pentoxide.

Phosphorus Pentoxide. It is obtained when phosphorus burns in a free supply of air or oxygen. It is a white. amorphous powder. It has no smell when pure. It has a strong affinity for water and is a hygroscopic substance. It is used as a drying agent for gases, dissolves in water with a hissing sound forming phosphoric acid :

Phosphorus pentoxide+water=Phosphoric acid

The affinity of phosphorous pentoxide for water is so great that it can remove the elements of water from many organic and inorganic substances, *e.g.*, it removes water from sulphuric acid and nitric acid leaving behind sulphur trioxide and nitric pentoxide :

Sulphuric acid − water = Sulphur trioxide.

Nitric acid − water = Nitric pentoxide.

Phosphoric acid (1) It is prepared by dissolving phosporous pentoxide in cold water :

Phosphorous pentoxide + water = Phosphoric acid.

(2) It is also prepared by the action of sulphuric acid on calcium phosphate (bone ash) :

Calcium phosphate + sulphuric acid = Calcium sulphate + phosphoric acid

(3) It can be prepared by boiling red phosphorus with nitric acid Nitric acid oxidizes phosphorus into phosphoric acid with the evolution of oxides of nitrogen .

Phosphorus + nitric acid = Phosphoric acid + oxides of nitrogen

Properties. It is a colourless crystalline solid. It is sold in the form of solution which has no smell but a strong acid taste. Its salts are known as phosphates, *e.g.*, when it acts upon caustic soda it forms sodium phosphate :

Caustic soda + phosphoric acid = Sodium phosphate + water.

Calcium phosphate is the most important salt as it is present in all fertile soils. Phosphates are used as fertilizers.

Matches. There are two kinds of matches in common use, *viz.*, the Lucifer matches and the Safety matches.

EXPERIMENT 108. (*a*) Grind together in a mortar equal quantities of potassium chlorate and sugar. Put the mixture on a slab and touch the mixture with a glass rod dipped in sulphuric acid. The whole mass begins to burn.

(*b*) Mix throughly in a mortar equal quantities of antimony sulphide, potassium chlorate and potassium dichromate. Put the mixture on slab and add a little red phosphorus. Stir the mixture. After a short time the mixture catches fire (The mixture will burn only when it is thoroughly mixed better with a hammer or spatula. Care should be taken for the burning particles fly apart.)

The first matches, called "dipping matches" were prepared by covering the end of the wooden splint with a paste of potassium chlorate, sugar and gum. This end was dipped in a bottle containing asbestos soaked in sulphuric acid. These matches did not become popular as the substances used are dangerous.

The 'friction matches' were manufactured by tipping the wooden splint with a mixture of antimony sulphide and potassium chlorate and were ignited by rubbing between the folds of a sand paper. The heat produced by friction causes the sulphur in the antimony sulphide to combine with the oxygen of the potassium chlorate

The Lucifer match or 'strike anywhere' matches are manufactured by dipping the splints of wood, in molten wax. They are tipped with a mixture of gum,

with phosphorus and potassium chalorate or red lead which is kept in a viscous form The tipped splints are then dried and packed in boxes. They can not be lighted by rubbing against any hard surface. In order to save them from the action of air and moisture, the tipped ends are dipped into varnish. When rubbed, the heat of the friction ignites the white phosphorus and brings about the combustion of wood.

White phosphorus, as already stated, is very poisonous. Its vapours produce a disease of the jaws, "Phossy Jaw" which was very common amongst the workers in these factories. Hence the use of phosphorus has been prohibited by law in some countries and phosphorus sulphide which is not poisonous and which requires greater friction is used In some factories scarlet phosphorus—a special variety of phosphorus—is used to avoid the poisonous effect of yellow phosphorus.

Safety matches now are manufactured on a large scale The splints are tipped with a paste of antimony sulphide, red lead, postassium chlorate, potassium dichromate and glue. They can ignite only when rubbed against the specially prepared surface attached to the box, which contains a mixture of red phosphorus, antimony sulphide, powdered glass and gum. When the match stick is rubbed against this surface a little of the red phosphorus ignites by the heat produced by friction and the combustion takes place.

SUMMARY

There are two varieties of phosphorus—white and red

The white phosphorus is now a-days prepared by heating together calcium phosphate, sand and coke in an electric furnace The vapours of phosphorus are collected and condensed

When white phosphorus is heated in closed vessels without contact of air to about 230 °C, red phosphorus is formed

White Phosphorus	Red Phosphorus
1 It is a white wax-like substance but turns yellow on exposure to light.	It is a violet red powder and is stable at the ordinary temperature
2. It is soft and can be cut with a knife	It is a powder
3 It is poisonous	It is non poisonous
4 Its density is 1 83	Its density is 2 1
5 It melts at 44° C, under water and burns at 34°	It does not melt but directly changes into vapour when heated to 290° C
6 It oxidizes very soon and glows in the dark--phosphorescent	It does not oxidize in the ordinary air and emits no light
7 It is kept under water to avoid its oxidation It should be handled under water	It need not be kept under water for it burns at 260 °C
8 It is soluble in carbon disulphide and other solvents	It is not soluble in carbon disuphide
9 It forms phosphorus pentoxide when burnt in air	It also forms phosphorus pentoxide when burnt in air

When phosphorus burns in a limited supply of air or oxygen phosphorus trioxide is formed which when dissolved in water forms phosphorus acid whose salts are known as phosphites

When phosphorus burns in an excess of air oxygen, phosphorus pentoxide is got which dissolved in water gives phosphoric acid whose salts are known as phosphates

Phosphates are used as fertilizers

White phosphorus is used in the manufacture of Lucifer matches Lucifer matches are tipped with a mixture of glue, white phospphorus and potassium chlorate or red lead and varnished They ignite when rubbed against a hard surface Phosphorus being poisonous has been replaced by phosphorus sulphide or an allotrope, scarlet phosphorus

Safety matches do not contain phosphorus in the tip The tip contains a mixture of glue, antimony sulphide, potassium dichromate and red lead. The special surface against which they are rubbed contains antimony sulphide, red phosphorus, Yellow phosphorus is used as a poison to kill rats

QUESTIONS

1 Give the occurrence of phosphorus Why does it not exist free in nature ?

2 Name the two forms of phosphorus. Give the characteristic difference between them How would you show their chemical identity ?

3 What happens when phosphorus burns in (*a*) a limited supply of air and (*b*) excess of air ? Give properties of each compound

4 How is phosphoric acid prepared ? What is it used for ?

5 How are phosphites and phosphates obtained ? What is the use of phosphates ?

6 How are lucifer and safety matches manufactured ?

Why is the use of white phosphorus disallowed ?

What is used in its place in the manufacture of matches ?

7 Distinguish between the properties of red and yellow phosphorus and mention what you know about phosphorus pentoxide

(*P-U , M.S.L C.*, 1925)

8 Write what you know about the different varieties of phosphorus and the characteristic differences between them How would you show their chemical identity ? What is the principal source of phosphorus and what are its important uses ? (*P U., M.S L C*,1931)

CHAPTER XII.

SODIUM CHLORIDE, HYDROCHLORIC ACID AND CHLORINE.

Sodium Chloride (*Common Salt*).

Occurrence. Everybody is familiar with this substance for it is an essential part of food. It is said that a man takes nearly 29lb. of salt in a year. It occurs abundantly and is very widely distributed in nature. Sea water contains about 3% of salt. The solid substances in the sea water contain about 77% of salt, 2% of potassium chloride, 10% of magnesium chloride, 10% magnesium and calcium sulphates, and a little of calcium bicarbonate and magnesium bromide. Most of the salt consumed in the world is obtained from sea water But in many countries, *e g*, in Austria and in the Panjab there are huge beds of rock salt of very great thickness. The salt range in Austria is said to be 500 miles long, 20 miles broad and 1200 feet thick. The Salt Range in the Panjab contains layers 1000 ft. thick. The salt in the Panjab is mined at Khewra, Kalabagh and Kohat. There is a salt range in the Mandi State also.

The rock salt is obtained by mining processes. In the Khewra mines there is a wall of translucent salt nearly 7 ft. thick Light burned on one side of the wall can be seen from the other.

The separation of salt from sea water. As stated above most of the salt in the world is obtained from

sea water or other salt lakes. In India the salt is obtained by the evaporation of sea water on the coasts of Bombay, Madras and Burma In Rajputana it is obtained from the Sambhar lake In these parts shallow beds are prepared near the sea coast above the level of the spring tides Sea water is carried to these beds and allowed to evaporate by the sun's rays. When the solution becomes concentrated enough, it is evaporated to dryness in iron pans by means of wood fires.

In cold countries, *e. g*, on the shores of the white Sea (Russia), the sea water is concentrated by freezing. Ice separates first, the remaining sea water is very rich in salt. It is further concentrated and salt got by evaporating over a fire.

Properties Common salt crystallizes in cubes. Rock salt when pure occurs in transparent or translucent cubic crystals, but often it is tinged yellow, brown pink, or even blue by impurities. The crystals have no water of crystallization and when a crystal of common salt is heated, it breaks with a crackling noise known as *Decrepitation*.

Salt is soluble in water. Its solubility in cold water and in hot water is practically the same. The solution is neutral to litmus Salt becomes moist in damp air due to the impurities in it of calcium and magnesium chlorides.

Uses. Salt is used for seasoning food. It is used for the preservation of meat, fish and *ghee*, etc. It is used for glazing earthenware, pottery, *e g.*, drain pipes, etc It is used in the alkali industry, for producing sodium carbonate and caustic soda. It is used in

freezing mixtures in some soap factories. In cold countries it is used to melt snow and ice on the roads. It is also used in the manufacture of compounds of chlorine.

Hydrogen Chloride (Hydrochloric Acid Gas)

Preparation. (1) Hydrochloric acid gas may be obtained by the direct combination of hydrogen and chlorine. Take a hard glass bottle and half fill it with hydrogen and half with chlorine. Place the flask containing the mixture of gases in sunshine. The gases combine chemically with an explosive noise and form hydrochloric acid gas.

(2) *Laboratory method* In the laboratory the gas is prepared by heating a mixture of common salt and strong sulphuric acid in a hard glass flask. The chemical change that takes place is

Sulphuric acid+common salt=Sodium sulphate

+hydrochloric acid gas.

Sodium sulphate is also called Glauber's Salt. This method is used to prepare the gas on an industrial scale as well.

EXPERIMENT 109 Fit up the apparatus as shown in the diagram Remove the stopper containing the thistle funnel and the delivery tube from the flask and put some powdered salt in the flask. Replace the stopper and pass the delivery tube into a cylinder. See that the apparatus is air-tight Add strong sulphuric acid in the flask through the thistle funnel, enough to cover the salt (Acid containing one part of water may as well be used) Heat the flask gently and collect the gas by the downward

displacement of air. In order to find out whether the jar is full or not, take a burning match stick or a moistened blue litmus paper near the mouth of the jar.

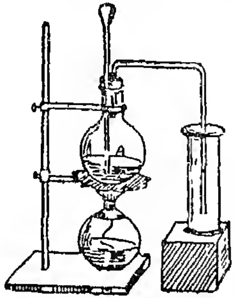


Fig. 56.

The match stick will be extinguished if the jar is full and the blue litmus will be turned red.

Collect a few jars with the gas to study its properties.

EXPERIMENT 110. (*a*) Examine the colour of the gas, smell it and taste it. It has no colour, but has a pungent smell and sour taste.

(*b*) Lower in the same jar a burning candle. Notice the gas does not burn and the candle goes out.

(*c*) In the same jar put a moistened blue litums, See that it is turned red.

(*d*) Invert a cylinder full of the the gas, removing the cover from the mouth. After a minute or so test the presence of gas by taking a burning match stick. There is no gas. It proves that the gas is heavier than the air.

(*e*) Remove the cover from the cylinder full of gas and blow on it Fumes are seen. This due to the combination of gas with water of the breath and formation of hydrochloric acid.

(*f*) Take a large, dry flask B, fitted with a rubber stopper carrying in it a tube drawn out into a jet. Fill the flask completely with hydrochloric acid gas. Invert the flask with the tube dipping into water coloured with blue litmus contained in a second large flask, A. as shown in fig. 57. Blow into the short tube in the lower flask to force a drop of water into the upper flask. The gas dissolves in that drop and water from the lower flask runs in the upper in the form of a fountain and turns red by the acid solution formed. (Fountain Experiment)

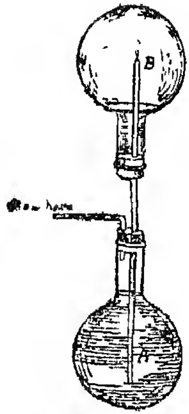


Fig 57.

A—blue litmus. solution.

B—fountain.

(*g*) Take a red dipped in ammonia near the mouth of jar full of gas and see the formation of dense white fumes In order to make the experiment more attractive proceed as below .—

EXPERIMENT 111. Take a cylinder and moisten its sides with hydrochloric acid. Take another cylinder and moisten its sides with ammonia. Place one above the other. Both will be filled with a dense white smoke. (*See* Fig 58.)

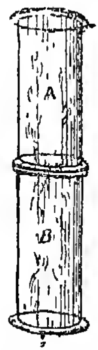


Fig 58.

A—hydrochloric acid.

B - ammonia.

Properties. Physical. It is a colourless gas with a very pungent irritating odour, which pinches the nose and throat. It has a sour taste. It is heavier than air and is extremely soluble in water, one volume of water at 0°C. and under normal pressure can dissolve 503 volumes of the gas.

Chemical. It does not burn or support combustion. It turns moistened blue litmus paper red—is acidic. It fumes with moist air forming hydrochloric acid. It forms dense white fumes with ammonia forming ammonium chloride:

Hydrochloric acid gas+ammonia
=Ammonium chloride.

The solution of the gas in water is called hydrochloric acid.

Composition. Hydrochloric acid gas is composed of equal volumes of chlorine and hydrogen. When one volume of chlorine and one volume of hydrogen are mixed and exposed in sunlight, they combine with an explosion to form two volumes of hydrochloric acid gas.

Hydrochloric Acid *Preparation.* It is prepared by dissolving hydrochloric acid gas in water

Hydrochloric acid gas+water=Hydrochloric acid.

Properties Physical. It is colourless when pure but commercial is generally yellow in colour owing to the presence of impurities. It fumes in moist air. It is sour in taste.

Chemical. It turns blue litmus red It gives strong fumes with ammonia. It stains wood or organic substances yellow. Many metals, such as zinc, iron,

etc., break up the acid-forming hydrogen and their chlorides

Zinc+hydrochloric acid=Zinc chloride+hydrogen.

Iron+hydrochloric acid=Iron chloride+hydrogen.

Similarly, magnesium, aluminium and tin are attacked by the cold dilute acid. Aluminium and tin are more readily attacked by strong hot acid. Mercury, silver, gold and platinum are not attacked either by strong or dilute acid, cold or hot. (Copper and lead are not dissolved by the cold dilute acid unless exposed to the air.)

The salts of this acid are known as chlorides, sodium chloride being the important.

Uses It is used in the manufacture of chlorine (for bleaching powder industry), in dyeing, calico printing, the manufacture of colours, in tanning industries and in cleaning metals. It is sometimes given as medicine and is present in the gastric juice of the stomach. It is an important reagent in the laboratory.

Test. It forms a white precipitate insoluble in acids with a solution of silver nitrate. The precipitate is soluble in ammonia and becomes black when exposed to sunlight:

Hydrochloric acid+silver nitrate=Silver chloride +nitric acid.

Silver chloride is insoluble and settles down as a precipitate. The same test holds good in the case of a soluble chloride, *e.g.*, sodium chloride.

EXPERIMENT 112 Dissolve a little of sodium chloride in a test tube. Add to it a few drops of silver nitrate solution. A thick white precipitate is formed :

Sodium chloride + silver nitrate = Silver chloride + sodium nitrate.

Note. Silver nitrate solution is kept in deep blue bottles for it turns black on exposure to sunlight. If put in ordinary bottle, it should be covered with black paper.

Aqua Regia. As already stated gold is not attacked by hydrochloric acid. However, when three volumes of hydrochloric acid and one volume of nitric acid are mixed, the mixture dissolves gold. It is known as aqua regia—Royal water, on account of its property of dissolving gold—'the king of metals.'

CHLORINE *Occurrence* Chlorine does not occur free in nature, but is found abundantly in the form of chlorides, chiefly sodium and magnesium chlorides

Preparation, (1) *From Hydrochloric acid* In the laboratory, chlorine is prepared by heating a mixture of manganese dioxide and concentrated hydrochloric acid in a flask The action takes place in two stages, *viz.*, the formation of a higher chloride (manganese chloride) and water in cold acid. When heated it decomposes into chlorine and leaves behind a lower chloride (manganous chloride) The complete reaction takes place as

Hydrochloric acid + manganese dioxide.

= Manganous chloride + chlorine + water.

EXPERIMENT 113 Fit up the apparatus as shown in fig 59 and be sure that it is air tight Remove the stopper fitted with thistle funnel and delivery tube and put some manganese dioxide in the flask Reinsert the stopper and add enough hydrochloric acid through the thistle funnel to cover the Manganese dioxide Warm gently, A greenish yellow gas is given out which soon fills the flask and the jar Remove the jar and keep it covered Collect a few jars of the gas to study its properties

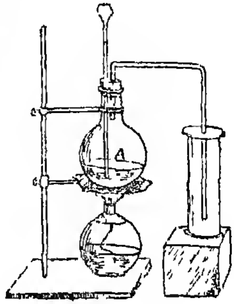


Fig 59

Note The gas is very poisonous and therefore, should be prepared in the open air or in a fume cupboard, if there is one in the laboratory After preparing the gas, the apparatus should be at once disconnected and washed All the experiments with the gas should be performed in the open air Do not inhale the gas

(2) *From Salt* Sometimes instead of using hydrochloric acid common salt and strong sulphuric acid are heated with manganese dioxide Here hydrochloric acid is produced first and is then oxidized by the manganese dioxide

(i) Sodium chloride + sulphuric acid + manganese dioxide =

Manganese dioxide + hydrochloric acid + sodium sulphate

(ii) Manganese dioxide + hydrochloric acid

= Manganous chloride + chlorine + water

The apparatus used for preparing gas by this method is the same as shown in fig,59

(3) Now-a days most of the chlorine used in industries is produced by the electrolysis of sodium chloride solution Chlorine is deposited at the anode and the sodium ions deposited on the kathode react with water to form caustic soda and hydrogen The action that takes place is —

Sodium chloride + water = Sodium hydrate (Caustic soda)

+ chlorine + hydrogen.

Hydrogen escapes at the negative pole

EXPERIMEMT 114 (*a*) Examine the colour of the gas It is greenish Its smell is very irritating It should not be smelled, it is poisonous

(*b*) Invert a cylinder full of the gas in a trough of water After a short time, some water rises into the jar This water has all the characteristics of the gas—chlorine water If the apparatus has not been disconnected pass the gas for a few minutes in a jar containing water to get good chlorine water

(*c*) Lower a burning candle in a jar full of the gas The candle continues burning dimly with a very smoky flame

(*d*) Put a moistened litmus paper, blue, or red The colour is bleached In the same jar put some moistened coloured flowers The colour is bleached In the same cylinder put a piece of paper writing your name on it See that the colour fades

Fig 60

(*e*) Warm a little turpentine oil in a dish and soak a filter paper in it Put the fitor paper in a cylinder full of the gas It catches fire giving out smoke

(*f*) Powder a piece of antimony in a mortar and sprinkle it in a jar of chlorine The particles burn brilliantly forming dense white fumes The fumes are of antimony chloride

(*g*) Take a small piece of phosphorus in a deflagrating spoon and lower it in a jar of the gas Phosphorus burns with a pale flame forming phosphorus chloride

(*h*) Put a small piece of sodium in a deflagrating spoon and heat it till it catches fire Introduce it into a jar of chlorine Sodium continues burning brightly forming sodium chlroide

(*i*) Throw a thin leaf of copper or Dutch Metal in a jar of chlorine The leaf at once catches fire

Properties (*a*) *Physical* Chlorine is a greenish yellow gas with a most irritating smell It is 2½ times heavier than air and is thus collected by the downward displacement of air It is fairly soluble in water, the water has all the peculiarities of the gas and is known as chlorine water

(*b*) *Chemical* It is as active as oxygen and combines readily with hydrogen to form hydrochloric acid If a jet of hydrogen is burned in a jar of this gas it burns with a greenish flame and fumes of hydrochloric acid are formed. It directly combines with most metals forming their chlorides, *e g*, antimony, copper, sodium, etc It combines with non-metals to form their chlorides *e g*, phosphorus forms phosphorus chloride when burnt in it Chlorine has a strong affinity for hydrogen and removes it from water, setting free oxygen

Water + chlorine = Hydrochloric acid + oxygen

This oxygen which is in the nascent condition bleaches vegetable colours Chlorine also removes hydrogen from carbohydrates and sets free carbon *e g* in the burning of candle and turpentine in the above experiments the smoke given out is the carbon set free Fumes of hydrochloric acid also accompany carbon smoke

Bleaching action of Chlorine Chlorine as stated above has strong affinity for hydrogen and removes it from water Thus when a substance whose colour is to be bleached is moistened and put in a jar full of chlorine, the gas takes up hydrogen from water forming hydrochloric acid and sets free oxygen The oxygen being in the nascent condition is very active and oxidizes the colouring matter into colourless oxide Dry chlorine cannot bleach colour

Bleaching Powder Bleaching powder is used for bleaching the colour of fabrics It is prepared by passing a stream of chlorine over dry slaked lime It is a mixture of calcium chloride and calcium hypochlorite It is a white powder and is more convenient to use than chlorine Its action as a bleaching agent depends upon the fact that it readily gives up its chlorine when acted upon even by weak acids like carbonic acid The chlorine liberated acts as stated above Bleaching powder is used for disinfecting purposes as well because chlorine acts upon organic substances, and destroys germs Bleaching powder is used for purifying water for drinking purposes in towns (*Chlorinated water*)

In order to use bleaching powder proceed as below

EXPERIMENT 115 Take two beakers In one prepare a solution of bleaching powder, in the other a very dilute solution of some acid say, sulphuric acid Dip the piece of cloth to be bleached first in the bleaching powder and then in the weak acid Repeat this process till the colour is completely bleached Finally wash the cloth in water to remove the excess of acid Final washing should be very thorough or the fabric will be damaged

Note Woollen and silken clothes are not bleached by this powder, as they are destroyed They are bleached by hydrogen peroxide sulphur dioxide

Uses of Chlorine It is chiefly used in bleaching cotton goods and paper pulp used in the manufacture of paper It is used as a disinfectant to purify water Large quantities of the gas are used in the manufacturing of bleaching powder

Chlorides. As stated above chlorides are the salts of hydrochloric acid. They may be prepared by the following methods :

(1) By the direct combination of chlorine with a metal, *e.g.*, copper or zinc :

Copper+chlorine=Copper chloride.

(2) By the action of hydrochloric acid on metals, *e.g.*,

Zinc+hydrochloric aicd=Zinc chloride
+hydrogen.

(3) By the action of hy ochloric acid on the oxides of metals, *e.g.*,

Zinc oxide+hydrochloric acid=Zinc
chloride+water.

(4) By the action of hydrochloric acid on the hydroxide of a metal, *e.g.*,

Sodium hydroxide+hydrochloric acid=Sodium
chloride+water.

(5) By the action of hydrochloric acid on the carbonate of a metal, *e.g.*,

Calcium carbonate+hydrochloric acid=Clacium
chloride+water+carbon dioxide.

Most of the chlorides are soluble in water. Silver chloride, mercurous chloride, cuprous chloride are however insoluble in water.

The test for a soluble chloride is the same as for hydrochloric acid.

SUMMARY.

Sodium chloride is very common in nature and is present in sea water, salt lakes and brine springs. Found in large beds as rock salt Rock salt is mined in the Panjab at Khewra and Kalabagh, etc. Salt from sea water is prepared by evaporating sea water in large beds

Sodium chloride is used in household as a preservative, in alkali industry and in the manufacture of compounds of chlorine

Hydrochloric acid gas. *Preparation* It is prepared by heating a mixture of sodium chloride and strong sulphuric acid in a flask an is collected by the upward displacement of air Hydrochloric acid gas is evolved and Glauber's salt is left in the flask

Properties It is a colourless gas with irritating odour It is sour to taste and is extremely soluble in water The solution is known as hydrochloric acid The gas does not support combustion nor is combustible It fumes in moist air and with ammonia forms ammonium chloride It turns moist blue litmus paper into red

Action of hydrochloric acid on metals may be summarized as below

Metals	Cold dilute Hydrochloric acid	Strong Hydrochloric acid
Zinc, magnesium and iron	Dissolve readily	Dissolve readily
Aluminium and tin	Dissolove slowly	Dissolve readily
Copper and Lead	Insoluble Are slowly attacked in the presence of air	Insoluble
Mercury Silver Gold, Platinum	Insoluble.	Insoluble

Uses. Hydrochloric acid is used in the manufacture of chlorine dyes, calico printing, manufacture of colours, in cleaning metals and in tanning industry. Used as a reagent in the laboratory.

Test. Silver nitrate gives a thick white precipitate of silver chloride which is soluble in ammonia.

Chlorine. It is prepared by heating (*i*) hydrochloric acid and manganese dioxide. (*ii*) A mixture of sulphuric acid and common salt is heated *with manganese dioxide*. (*iii*) By the electrolysis of a chloride, chiefly sodium chloride.

Properties. It is greenish-yellow gas with a disagreeable irritating smell. It is heavier than air and soluble in water. The solution is known as chlorine water. It is non-combustible but supports combustion. Phosphorus and sodium burn in it to form their chlorides. Candle burns in it with a sooty flame. It is very active. It has a great affinity for hydrogen, removes it from water, thus setting free oxygen which oxides and bleaches vegetable colours.

Uses. It is used as a bleaching agent and disinfectant. It is used in the manufacture of bleaching powder.

Sulphur dioxide	*Chlorine*
Sulphur dioxide bleaches vegetable colours in the presence of moisture.	Chlorine bleaches colours in the presence of moisture.
Sulphur dioxide takes up oxygen from the moisture and sets free hydrogen which being in the nascent condition is very active and reduces the colouring matter.	Chlorine removes hydrogen from moisture and sets free oxygen which being in the nascent condition is very active and oxidizes the colouring matter.
The bleached colour can be restored on exposure to air or by the addition of a few drops of sulphuric acid.	Colour once bleached cannot be restored.
It does not injure delicate fibres like wool and silk.	It destroys silk and wool fibres.

QUESTIONS.

1. Give the occurrence of common salt. How is it obtained from sea water? What are its uses?

2. How is hydrochloric acid gas prepared in the laboratory? Give its properties and uses.

3. Describe the action of hydrochloric acid on important metals. How are chlorides obtained? What is the test for a chloride or hydrochloric acid?

4. What happens when sodium chloride solution is electrolysed

5. How is hydrochloric acid obtained? Give a short account of its chief properties and uses? (*P U*, *M* & *S L C*, 1925)

6. Describe carefully the method of preparation and collection of hydrochloric acid gas. What experiments will you perform to show its properties? What is its composition? How can it be determined by the synthetic method? (*P U*, *M* & *S L.C*, 1917)

7. Describe the chemical changes that take place when dilute hydrochloric acid acts upon (1) Silver nitrate, (2) Iron, (3) Ammonia (4) Caustic soda, (5) Silver and (6) Magnesium

8. You are provided with four bottles, containing water, dilute sulphuric acid, dilute hydrochloric acid and soda water. What experiments will you perform to distinguish them from one another?

9. Describe the preparation, properties and uses of chlorine

10. Describe what happens when each of the following is introduced in a jar of chlorine. Give reasons

(*a*) A burning candle, (*b*) a filter paper soaked in hot turpentine, (*c*) a moistened flower, (*d*) phosphorus, (*e*) powdered antimony, (*f*) copper leaf

11. A piece of coloured cloth is dropped in a jar of dry chlorine. Will the colour bleach? What will you do to bleach the colour and why?

12. How is bleaching powder prepared and used? Why is it preferred to chlorine?

13. Compare the bleaching action of chlorine and sulphur dioxide

14. Write short notes on (*a*) Decrepitation (*b*) Chlorine water, (*c*) Disinfecting property of chlorine

15. Account for the fumes seen when (*a*) hydrochloric acid gas comes in contact with moist air, (*b*) when a rod dipped in ammonia is brought near hydrochloric acid

16. What happens when—

(1) Sulphuric acid is heated with common salt?

(2) Sulphuric acid, common salt, and manganese dioxide are heated together?

(3) Soda water is added to the bleaching powder?

(4) Equal quantities of chlorine and hydrogen are mixed and exposed to sunlight?

(5) Three volumes of hydrochloric acid are mixed with one volume of nitric acid? What is this mixture used for?

CHAPTER XIII.

COMPOUNDS OF NITROGEN.

Ammonia *Occurrence.* It is the most important compound of nitrogen and hydrogen. It occurs in small traces in the atmosphere and in certain waters. Its presence in these cases is due to the decomposition of organic substances containing nitrogen. The bad smell in urinals and in the stables is due to this gas. If water smells of ammonia, it should not be used for household or drinking purposes

In the combined state ammonia exists in the soil as ammonium nitrate and ammonium chloride (*naushadar*) Ammonium chloride and ammonium sulphate are found deposited on the sides of the craters of active volcanoes.

Ammonia was known to ancients as 'the spirit of hart's horns' as it was prepared by the destructive distillation of hoofs, horns, hair, hides, etc , of animals

Preparation. Ammonia is prepared in the laboratory by either heating concentrated solution in a flask or most commonly by heating a mixture of ammonium chloride and dry slaked lime in a flask. The chemical action takes place as follows:—

Ammonium chloride+calcium hydroxide = Calcium chloride+ water+ammonia.

EXPERIMENT 116. Fit up the apparatus as shown in the Fig.61. Prepare a mixture of equal quantities of ammonium chloride and dry slaked lime. Put the mixture in the flask and fit it with the stopper containing the the delivery tube as shown in the diagram. See that the apparatus is air-tight Heat the flask

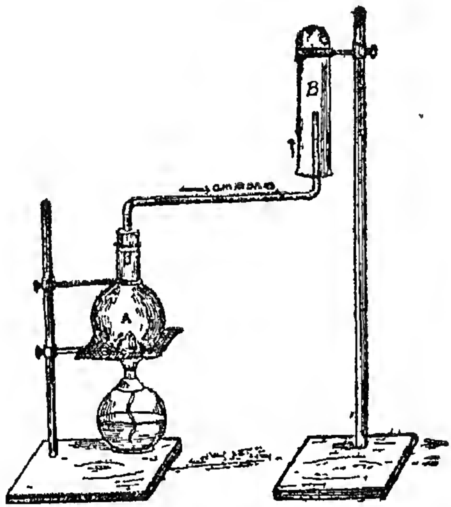


Fig 61

A—Flask containing ammonium chloride and slaked lime B—Ammonia Gas

gently and collect the gas by the upward displacement of air. If dry ammonia is requied it is passed through quick lime contained in a vessel. Collect a few cylinders of the gas and keep them aside, with their mouths covered. Perform the following experiments to study the properties of the gas.

EXPERIMENT 117. Notice the colour. (*a*) Remove the disc from a jar and smell the gas. It brings tears in the eyes

(*b*) Put in the cylinder a moist red litmus paper. It is turned blue.

(*c*) Bring a rod dipped in hydrochloric acid near the mouth of the jar. Notice the formation of dense fumes.

(*d*) Take a rod dipped in nitric acid near the mouth of a jar full of the gas. Notice the formation of white fumes.

(*e*) Invert a jar full of the gas in a trough of water The water soon rises and fills the jar.

(*f*) Perform the fountain experiment as in the case of hydrochloric acid—Exp.110 (*f*). In the lower

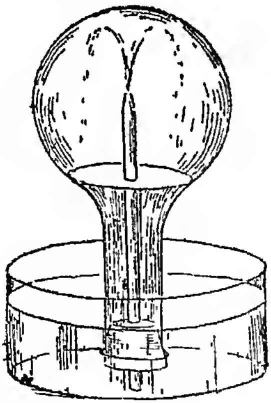


Fig 62.

flask put water coloured with red litmus. It will be turned blue.

For the fountain experiment apparatus shown in Fig. 62 may be used. It is simpler than the other

(g) Take a paper dipped in turmeric solution and put it into a jar of the gas. It is turned reddish brown

(h) Dissolve in a test tube some copper sulphate, add to the solution some liquid ammonia. A deep blue colour is produced. This is used as a test for ammonia. Any salt of copper may be used

Properties. Physical It is a colourless gas with a characteristic pungent smell When inhaled suddenly it brings tears in the eyes It is lighter than air and is, therefore, collected by the upward displacement of air It is extremely soluble in water as is illustrated by the fountain experiment above. It has a burning taste. It is easily liquefied by cold or pressure, to a colourless liquid which is used in refrigerating machines (ice manufacturing machines) and is sold in large cylinders

Chemical It does not burn nor supports combustion. It, however, burns in an atmosphere of oxygen with a greenish-yellow flame The flame is in reality due to the burning of hydrogen set free when ammonia is heated Thus we get nitrogen and water when it burns in oxygen It turns red litmus blue and yellow turmeric brownish red It fumes with hydrochloric acid and nitric acid, forming ammonium chloride and ammonium nitrate respectively. With sulphuric acid, ammonium sulphate is formed. Thus ammonia combines directly with acids forming their respective salts.

It does not fume with sulphuric acid When it dissolves in water, heat is given out and ammonium hydrate, a strong alkali, is formed.

In the presence of heated platinum as a catalyst, a mixture of ammonia with air or oxygen is oxidized to nitric acid :

Ammonia + oxygen = Nitric acid.

Manufacture On a large scale ammonia is prepared by heating the ammoniacal liquor got in the distillation of coal The liquor is mixed with lime and heated to get the gas

In most countries it is prepared by synthetic methods, *i e.*, combining nitrogen and hydrogen under certain conditions.

Ammonia is made up of one part of nitrogen and three parts of hydrogen by volume.

Uses Ammonia gas is used in the manufacture of sodium carbonate (washing soda). Liquid ammonia is used in refrigerating machines. Its use in these machines is due to the fact that liquid ammonia absorbs heat in evaporating. It absorbs 260 cals. per gram, *i.e.*, one gram will be able to freeze $\frac{260}{80}=3$ grams of water at 0° into ice. It dissolves grease and hence is used as a cleansing agent Its salts are used in medicine, *e.g.*, ammonium carbonate is sold as smelling salt. Ammonia liquor is used in case of wasp sting. Ammonium hydrate—the solution of ammonia in water—is used as a reagent in the laboratory.

Test for Ammonia. The gas is detected by its smell; by its action on turmeric paper, turning it reddish brown and by producing fumes with hydrochloric acid. Liquid ammonia produces a deep blue colour with copper salts.

Nitric acid

Occurrence. Nitric acid is another very important compound of nitrogen with hydrogen and oxygen. It occurs freely in small traces in the atmosphere. It is widely distributed in nature in the form of nitrates, chiefly sodium nitrate (Chili saltpetre) and potassium

nitrate or nitre. Nitre is common in the Punjab, chiefly in Shahpur Mianwali and Rohtak districts where there are big works for refining nitre.

It was known to the ancients under the name of aqua fortis (strong water) probably due to its property of dissolving most of the metals.

Preparation. In the laboratory it is prepared by distilling a mixture of potassium nitrate with concentrated sulphuric acid in a glass retort and collecting the acid in a receiver cooled by water. The action that takes place is :

Potassium nitrate+sulphuric acid=Potassium sulphate+nitric acid.

Intsead of potassium nitrate, sodium nitrate may be used.

EXPERIMENT 118. Fit up the apparatus as shown in the Fig. 63. In the retort put a small quantiry of potassium nitrate and enough concentrated

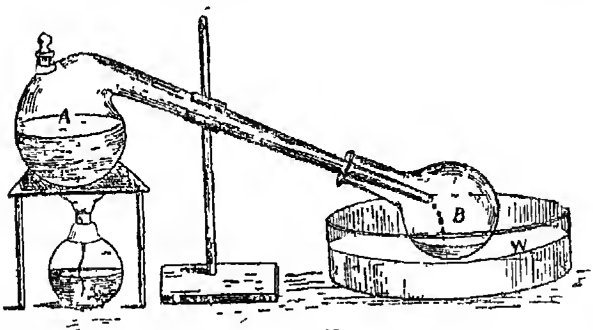


Fig. 63.

sulphuric acid to cover it. Fix it on the ring of the clamp stand and connect the neck with a flask kept in cold water, as in the diagram 63. Plug the the mouth of the flask by cotton so that fumes may not come out.

Heat the retort gently. Brown fumes are given off in abundance and soon drops of a light yellow liquid are seen to fall in the reciever. Collect sufficient amount of the acid. Perform the following experiments to study its properties.

EXPERIMENT 119 (*a*). See the colour of the liquid, smell it. Take a drop of the liquid and put it in a glass of water. Take it. It is sour to taste.

(*b*) Take a little of the acid in a test tube and dilute it with water. Put in it small pieces of blue and red litmus paper. Blue litmus is turned red

(*c*) Heat a few drops of nitric acid in a test tube, a reddish gas is given off.

(*d*) Heat strongly a quantity of saw dust in an iron tray till it is charred. Add a few drops of strong nitric acid on it. It begins to burn.

In (*c*) and (*d*) use nitric acid prepared in the laboratory to get good result.

(*e*) Repeat experiment (*d*) with a red-hot glowing charcoal. It bursts into a flame.

(*f*) Fix a clay tabacco pipe with a long item in a clamp stand as shown in the Fig 64. Connect the stem with a rubber tube which leads into a trough of water. Invert over it a test tube full of water. Heat strongly the stem of the pipe. Add a small quantity of concentrated nitric acid in the bowl. Bubbles of a gas are seen to rise in the tube. Remove the tube when full of the gas. Test the gas by introducing in it a glowing splinter. It begins

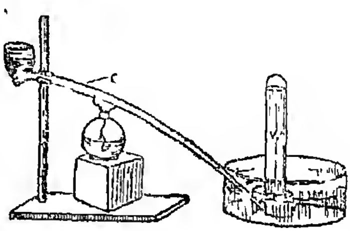


Fig 64

to burn showing thereby that the gas is oxygen. Nitric acid when heated gives up its oxygen. It is an oxidizing agent

(*g*) Test the action of acid, both strong and dilute, on zinc, copper, lead, iron, and magnesium.

(*h*) Introduce a match stick in the acid. It is turned brown.

(*i*) Take some nitric acid in a test tube and add to it gently a solution of ferrous sulphate (iron sulphate). On the surface is formed a dark ring which disappears on shaking.

Repeat the same experiment with dilute nitric acid or a solution of some nitrate. Before adding ferrous sulphate add in this case a few drops of strong sulphuric acid. The tube becomes hot Cool the tube and then add gently the sulphate solution. The black ring appears as in the experiment (*i*) and it disappears on shaking. This experiment is used as a test for nitric acid or nitrates The sulphuric acid being heavier settles at the bottom of the test tube and the black ring appears at the junction of the acid and the solution.

Properties Physical. Nitric acid is a colourless liquid when pure Commercial acid is pale yellow in colour due to nitrogen peroxide dissolved in it. Its specific gravity is 1·52 and boiling point is 86°C

Chemical. It fumes in air It is very corrosive, and causes painful wounds. It stains skin deep yellow. A piece of wood is turned brown in it. It turns blue litmus red. It decomposes on heating, giving off oxygen and oxides of nitrogen. It is an oxidizing agent. Glowing charcoal or heated saw dust burns when a few drops of this acid are added because of the oxygen

which it gives up readily to support combustion. Phosphorus is oxidized to phosphoric acid and sulphur to sulphuric acid by concentrated nitric acid :

Phosphorus + nitric acid = Phosphoric acid + oxides of nitrogen.

Sulphur + nitric acid = Sulphuric acid + oxides of nitrogen.

It acts upon most of the metals giving off oxides of nitrogen and forming nitrates of the metals. Magnesium is the only metal which liberates hydrogen when dissolved in dilute nitric acid.

Ferrous sulphate reduces nitric acid to nitric oxide which dissolves in excess of ferrous sulphate solution giving a brown ring.

Action on metals. Sodium, copper, silver, zinc, mercury, iron and lead when acted upon by nitric acid form their respective nitrates, water and oxides of nitrogen. Aluminium, gold and platinum are not attacked by the acid, dilute or strong.

Tin with strong nitric acid gives metastannic acid, water and oxides of nitrogen. In dilute nitric acid it slowly dissolves to form tin nitrate, ammonium nitrate and water.

Magnesium is the only metal which displaces hydrogen from dilute nitric acid. (Use very dilute nitric acid.)

Manufacture. On a large scale the acid is prepared by the distillation of sodium nitrate mixed with strong sulphuric acid in large cast iron retorts, and the nitric acid is condensed in stoneware or earthenware vessels kept in cold water. Much nitric acid is now made by the arc process from the atmospheric nitrogen (the

fixation of nitrogen) and by the oxidation of ammonia.

Uses of Nitric Acid. It is used as a solvent for metals, and for etching on copper. It is used widely in the manufacture of dye-stuffs, explosives and sulphuric acid. It is used in batteries. Its salts are used in calico printing, photography, etc.

Nitrates. Nitrates are the salts of nitric acid. They are obtained in the following ways :—

1. By the action of metals on nitric acid, *e.g.*,
   Copper + nitric acid = Copper nitrate + oxides of nitrogen + water.

2. By the action of nitric acid on the oxide or hydroxide of a metal, *e.g.*,

   Nitric acid + Lead oxide = Lead nitrate + water
   Nitric acid + sodium hydroxide = Sodium nitrate + water.

3. By the action of nitric acid on carbonates, *e.g.*
   Sodium carbonate + nitric acid = Sodium nitrate + carbon dioxide + water

4. Ammonia combines directly with nitric acid to form ammonium nitrate.

All nitrates are soluble in water. When heated strongly all decompose. Nitrates of heavy metals, *e.g.*, lead, iron, copper, etc., when heated decompose into oxides of nitrogen and oxygen leaving behind an oxide of the metal :

Lead nitrate → Lead oxide + oxygen + oxide of nitrogen.

Nitrates of alakli metals, *e.g*, sodium and potassium give out oxygen leaving a nitrite, thus :

Potassium nitrate ⟶ Potassium nitrite + oxygen

Ammonium nitrate when heated gives water and a gas called nitrous oxide.

Nitrates act as powerful oxidizing agents when heated.

EXPERIMENT 120. Heat in a test tube a small quantity of Potassium nitrate till it melts, and then drop into it a piece of dry charcoal. The charcoal ignites and burns violently forming carbon dioxide.

Similarly sulphur is oxidized to form sulphur dioxide when heated with a nitrate.

Uses. Nitrates are used in photograpy, calico printing and as fertilizers. They are used in explosives and in fireworks. They are used in the manufacture of nitric acid.

Test of Nitric acid and Nitrates :

Nitric acid. (1) Copper test—Nitric acid, dilute or strong, dissolves copper with the evolution of red oxides of nitrogen.

(2) Nitric acid forms a brown ring when ferrous sulphate solution is added to it.

Nitrate. (1) When heated they give up red oxides of nitrogen or the gas oxygen

(2) Solution of a nitrate to which a few drops of sulphuric acid are added gives a brown ring with ferrous sulphate solution.

(3) Copper and strong sulphuric acid when heated with a nitrate give red fumes of oxides of nitogen.

Fixation of Nitrogen. Nitrogen is a very inert element, *i.e.*, it does not combine with other elements. It has been, however, discovered that the atmospheric nitrogen can be fixed, that is, made to combine with other elements. If electric sparks are passed through an enclosed volume of air, about 1% of nitrogen is converted into a brown gas– nitric oxide This gas is cooled and oxidized into nitrogen peroxide which is dissoloved in water to form nitric acid. From nitric acid the nitrates which are required as fertilizers or for other industries are obtained. This industry is chiefly carried on in Norway, where electric supply is very cheap. The process is known as Arc process

Another method to make use of the atmospheric nitrogen is to fix it with hydrogen to form ammonia Nitrogen is obtained by the liquefaction of air and hydrogen from the electrolysis of water or from water gas. The mixture of these two gases is subjected to electric sparks. Some ammonia is formed. This ammonia is either converted into ammonium sulphate, a fertilizer, or it may be oxidized to form oxides of nitrogen to get nitric acid and nitrates. Most of the fixed nitrogen is obtained by this way.

Thus the atmospheric nitrogen has become a never-ending source of nitric acid and its compounds and the danger of natural nitrates in the world being used up has been removed.

SUMMARY

The most important compounds of nitrogen are ammonia (nitrogen and hydrogen) and nitric acid (nitrogen with hydrogen and oxygen) Ammonia is prepared in the laboratory by heating together a mixture of ammonium chloride and dry slaked lime in a flask The gas is collected by the downward displacement of air.

Properties It is a colourless gas with pungent tear bringing smell and is burning to taste It is lighter than air, and is extremely soluble in water It does not support combustion but burns in an atmosphere of oxygen or in air if a flame is constantly applied to it It is an alkaline gas and has all the properties of an alkali It combines directly with acids to form salts ammonium sulphate, chloride or nitrate

Uses. Liquid ammonia when evaporated absorbs much heat and is therefore used in the freezing machines It is used in the manufacture of washing soda and nitric acid Its salts are used as smelling salts

Test Turns red litmus blue, yellow turmeric brown and fumes with hydrochloric acid

Nitric acid is prepared and manufactured by distilling a mixture of sulphuric acid and sodium nitrate The distillate is nitric acid

Properties It is a colourless liquid when pure, commercial acid is yellow in colour It has all the properties of an acid, i e , turns blue litmus red and is sour to taste It is 1 52 times heavier than water and its boiling point is 86°C

It attacks most of of the metals forming their nitrates It is an oxidizing agent and decomposes into oxygen and oxides of nitrogen when heated It fumes in moist air with ammonia Magnesium is the only metal which displaces hydrogen from dilute nitric acid Aluminium gold and platinum are the only metals which are not attacked by it Gold and platinum are, however attacked by aqua regia

Uses It is used in many industries, e g dye-stuffs explosives, sulphuric acid, etc. Its salts are used in photography, calico printing, and as fertilizers It is used for etching on copper vessels,

Test—Ring test Add to the strong acid a few drops of ferrous sulphate, a brown ring is formed In case of dilute acid add a few drops of sulphuric acid before the addition of sulphate The same test is used for nitrates

Nitartes are the salts of nitric acid with metals. They may be obtained by the action of acid on metals, their oxides or hydroxides or carbonates

Nitrates are used as fertilizers and in photography or calico printing and dye-stuffs.

Atmospheric nitrogen has been made to combine with oxygen or hydrogen to form nitric oxide or ammonia This is known as fixation of nitrogen.

QUESTIONS.

1 How is ammonia prepared ? Give a short account of its chief properties and uses ? *(P. U , M & S L C , 1917)*

2 State the composition and properties of ammonia What happens when it is mixed with (a) water, and (b) nitric acid, (c) hydrochloric acid ?

3 You are provided with some cylinders of ammonia gas Describe the experiments you would perform to show its physical and chemical properties.

4 How is nitric acid prepared in the laboratory ? Give its physical and chemical properties To what uses is it put ?

5 You are provided with four bottles containing solutions of carbon dioxide, sulphuric acid, hydrochloric and nitric acid in water What experiments would you perform to identify one from the other ?

6. Describe the preparation of nitric acid and give experiments to show the oxidizing properties of the acid and its salts

(P U, M & S L C, 1948)

7 How is ammonia prepared on a large scale ? What are its uses in commerce ?

8. Describe an experiment to show ammonia is extremely soluble in water. What is the name of the resulting substance ? Prove that it is alkaline in nature

9 You are provided with four bottles containing water and solutions of sulphuric acid, hydrochloric acid and ammonia What test would you try to distinguish one from the other ?

10 What is the action of

(a) nitric acid on slaked lime ?

(b) ammonium chloride on caustic soda ?

(c) nitric acid on heated charcoal ?

(d) nitric acid on magnesium ?

11 What are nitrates ? How are they prepared ? What is their use ?

12 What happens when

(a) lead nitrate is heated ?
(b) potassium nitrate is heated ?
(c) sodium nitrate and sulphur or charcoal are heated together ?
(d) sulphur or phosphrus are heated with nitric acid ?
(e) ammonium nitrate is heated

13. What test will you make to distinguish a nitrate ?

14 Describe the action of important metals on nitric acid

15. What do you understand by the process 'Fixation of nitrogen' ?

CHAPTER XIV

ACIDS, BASES AND SALTS

Compounds can be grouped into three classes according to their behaviour towards litmus :—

Acids—which turn blue litmus red

Bases or alkalies—which turn red litmus blue

Salts which have no effect upon litmus.

Acids We have already seen that acidic oxides when dissolved in water gave acids. The most important acids are :—

1 Sulphuric acid Sulphur, oxygen and hydrogen.

2. Hydrochloric acid Hydrogen and chlorine.

3 Nitric acid - Hydrogen, nitrogen and oxygen

If we study the composition of all these acids we find that the element hydrogen is present in all these and that oxygen is not an essential constituent. At one time it was thought that acids contained oxygen but later on such acids, *e.g.*, hydrochloric acid, hydriodic acid, etc., were found which did not contain oxygen. However, it is found that all the acids contain hydrogen as a necessary constituent and not oxygen.

General methods of forming acid may be summarized as—

1. Direct union of water and acidic oxides or acid anhydrides as they are called, *e.g.*,

(i) Sulphur trioxide+water=Sulphuric acid.

(ii) Carbon dioxide + water = Carbonic acid

(iii) Phosphorus pentoxide + water = Phosphoric acid.

2 The salt of the acid which is to be prepared is taken and treated with a strong acid, generally sulphuric acid or hydrochloric acid :

(i) Sodium chloride + sulphuric acid = Sodium sulphate + hydrochloric acid gas.

(ii) Potassium nitrate + sulphuric acid = Potassium sulphate + nitric acid.

(iii) Sodium carbonate + hydrochloric acid = Sodium chloride + carbon dioxide + water

General properties of acids.—

1 The acids have a sour taste.

EXPERIMENT 121. Take three beakers full of water. Put two drops of sulphuric acid in one, hydrochloric acid in another and nitric acid in the third. Taste the weak solutions All are sour to taste

In reality the sour taste of a substance is due to the presence of some acid in it. Raw fruits are sour due to the presence of acids The pickles are sour due to the same reason.

2 Acids turn blue litmus red

EXPERIMENT 122. Prepare a weak solution of hydrochloric acid in a test tube. Put blue litmus paper in it It is at once turned red. Test your pickles and vinegar with blue litmus. They turn blue litmus red. Urine turns blue litmus red because it has an acid in it.

3. All acids contain hydrogen as an essential constituent.

It has been already studied above that hydrogen and not oxygen is a necessary part of each acid.

4. The hydrogen of the acids can be displaced by certain metals.

EXPERIMENT 123.(*a*) Take some dilute sulphuric acid in a test tube. Put iron nails in it. Invert over it a test tube for some time. Bring a burning taper near the mouth of the tube. The gas burns with explosion. This is hydrogen.

(*b*) Repeat the same experiment with dilute hydrochloric acid. Hydrogen is given off in this case as well.

(*c*) Take dilute nitric acid in a test tube and put a piece of magnesium in it. Effervescence takes place and a gas is given out which can be tested to be hydrogen.

The above three experiments prove that hydrogen is present in all acids and that it can be displaced by metals

5. Acids decompose carbonates.

EXPERIMENT 124. Take sodium carbonate (washing soda) in a test tube and add to it dilute hydrochloric acid. A brisk effervescence takes place and carbon dioxide is given out. The gas can be distinguished by passing it through lime water.

Any carbonate, *e.g.*, marble, chalk, *cowries*, etc., may be taken instead of washing soda and any acid substituted In each case carbon dioxide is evolved.

(*b*) Acids react with bases and basic oxides to form salts.

EXPERIMENT 125. In a beaker take a weak solution of hydrochloric acid in water Add a few drops of blue litmus solution to it. It turns red. Now add to it drop by drop a weak solution of caustic soda till the red colour of the acid just disappears. Test it with blue and red litmus paper. Both remain unaffected. The liquid is neutral to litmus. Evaporate it on a sand bath. A residue is left behind Taste it and satisfy yourself that it is common salt.

Similarly, if we take any acid and any alkali and treat as above we get a salt

We are now in a position to define the term acid more correctly as we have studied the general properties of acids. *An acid is a substance which has sour taste, turns blue litmus red and which has hydrogen as an essential constituent which can be replaced by a metal*

Bases and Alkalies Bases are generally the oxides or hydroxides of metals. They are formed when a basic oxide is dissolved in water Those bases which are readily soluble in water have a soapy touch and turn red litmus blue, are termed as *Alkalies.* The most important alkalies we have studied so far are :—

(1) Sodium hydroxide or caustic soda=sodium, hydrogen and oxygen.

(2) Potassium hydroxide or caustic potash=potassium, hydrogen and oxygen

(3) Calcium hydroxide or lime water=calcium, hydrogen and oxygen.

(4) Ammonium hydroxide=nitrogen, hydrogen and oxygen

General methods of the preparation of Alkalies —

1. Direct union of a basic oxide with water, *e.g.*, Calcium oxide+water=Calcium hydroxide

2. Reaction of a metal with water, *e.g.*, Sodium +water=Sodium hydroxide+hydrogen

3. Electrolysis of salt, *e.g.*, Sodium chloride+ water=Sodium hydroxide+hydrogen+chlorine.

4. Double decomposition of salts with other bases, *e.g.*, Sodium carbonate+clacium hydroxide= Sodium hydroxide+calcium carbonate.

General properties of Alkalies —

(1) They contain both hydrogen and oxygen. This is clear if we study the domposition of the obove-mentioned alkalies.

(2) Alaklies turn red litmus blue.

EXPERIMENT 126 Dissolve caustic soda in a beaker containing water Put red litmus solution in it. It at once turns blue. Repeat the same experiment with other alkalies. Place a red litmus paper in your mouth. It turns blue. This proves that the sputum is alkaline

(3) Alkalies have a soapy touch and bitter taste

EXPERIMENT 127 Dissolve some caustic potash in water Dip your finger in the solution and rub the fingers. A soapy touch is felt. Taste a drop, it is very bitter. Rinse your mouth after tasting.

(4) Alaklies turn yellow turmeric brown

Prepare turmeric solution in a test tube. Add to it a few drops of the solution of caustic soda. The yellow colour is at once turned into reddish-brown.

You might have noticed this action when a drop of soup falls on your clothes and you wash it with soap. The colour becomes reddish brown. The soap contains either caustic soda or caustic potash. They are alkalies and bring about the above change

(5) Alaklies turn phenolphthalein red.

EXPERIMENT 128. Prepare a solution of phenolphthalein in a test tube Add to it a few drops of caustic soda. A deep red colour is obtained.

(6) Alaklies react with acids or acidic oxides to form salts.

This has already been seen while studying the general characteristics of acids.

A base then may be defined as *an oxide or a hydroxide of a metal which is capable of neutralizing an acid to form a salt and water.* It contains both oxygen and hydrogen as its essential constituents

Salts Salts are compounds formed by the action of an acid on a base or alkali and *vice versa.* Salts are neutral to litmus. The important salts are sulphates, chlorides, nitrates, and carbonates which are formed by the neutralization of sulphuric acid, hydrochloric acid, nitric acid and carbonic acid, respectively, by some alkali

Formation of Salts. The important methods which we have read so far are:—

(1) By the neutralization of an acid by a base. When an acid and alkali are brought together they lose their characteristic properties and are said to neutralize each other. The process is known as *neutralization.*

EXPERIMENT 129. Neutralize hydrochloric acid, sulphuric acid, nitric acid, respectively, by caustic soda. Evaporate the neutral liquids and get the salts .

Hydrochloric acid+caustic soda=Sodium
chloride+water

Sulphuric acid+caustic soda=Sodium
sulphate+water.

Nitric acid+caustic soda=Sodium nitrate+water.

This process is very useful as it gives an antidote for acid poisoning. If one drinks acid give him a weak alkali and if one takes alakli, give him a weak acid, *e.g* , vinegar or lime-juice

(2) By the action of an acid on a metal

EXPERIMENT 130. Put some zinc in dilute sulphuric acid. When effervescence stops, pour off the liquid. Evaporate it. The residue is a salt, zinc sulphate :

Zinc+sulphuric acid=Zinc sulphate+hydrogen.

(3) By the action of an acid on the oxide or hydroxide af a metal ;

(*a*) Zinc oxide+sulphuric acid=Zinc sulphate
+water.

Lead oxide+nitric acid=Lead nitrate+water

(*b*) Carbonic acid+lime water (calcium
hydroxide)=Calcium carbonate+water.

(4) By the action of an acid on a salt of another acid.

Sodium chloride+sulphuric acid=Sodium
sulphate+hydrochloric acid gas.

(5) By double decomposition, *e.g.*, when silver nitrate and sodium chloride react they form two new salts, silver chloride and sodium nitrate :

Silver nitrate+sodium chloride=Silver chloride +sodium nitrate.

(6) By the direct union of a metal and a non-metal, *e.g.*, when antimony is thrown in chlorine we get antimony chloride. Again zinc and sulphur heated together give zinc sulphide and so on :

Antimony+chlorine=Antimony chloride.
Zinc+sulphur=Zinc sulphide.

(7) By the interaction of a metal with a salt of another metal, *e.g.*, when iron is thrown in copper sulphate, iron displaces copper and forms iron sulphate :

Iron+copper sulphate=Iron sulphate+copper.

Naming of salts. We have seen that the compounds generally end either in 'ide,' or 'ite' or 'ate'. The compounds which contain two elements end in 'ide', *e.g.*, zinc chloride contains zinc and chlorine, lead oxide contains lead and oxygen; iron sulphide contains iron and sulphur; ammonia contains nitrogen and hydrogen=hydrogen nitride; water, hydrogen and oxygen=hydrogen oxide; hydrochloric acid hydrogen and chlorine=hydrogen chloride, and so on.

Those compounds which end in 'ite' contain some oxygen and are generally salts of acids which end with 'ous'; *e.g.*, sodium nitrite is a compound of nitrous acid, and sodium sulphite is a salt of sulphurous acid. Similary, phosphites are salts of phosphorous **acid.**

The compounds which end in 'ate' contain a sufficient amount of oxygen, a part of which can be

replaced, *e.g.*, sodium sulphate contains sodium, sulphur and oxygen; potassium nitrate contains potassium, nitrogen and oxygen. They are generally the salts of acids which end in 'ic', *e.g.*, sulphates are salts of sulphuric acid, nitrates, of nitric acid and phosphates, of phosphoric acid.

Potassium chlorate when heated parts with the whole of its oxygen and leaves behind potassium chloride

SUMMARY,

An acid is a chemical compound with sour taste, which turns blue litmus red and which has got hydrogen as an essential constituent which can be replaced by a metal

The General Properties of Acids They are sour, turn red litmus blue, they have hydrogen as a necessary part, they decompose carbonates, they give hydrogen when acted upon by metals and they neutralize bases to form salts

Acids are generally obtained by dissolving acid anhydrides in water

A base is an oxide or a hydroxide of a metal and is capable of neutralizing an acid to form salt and water. It contains oxygen and hydrogen as its necessary part

The General Characteristics of Bases They turn red litmus blue, yellow turmeric brown, phenolphthalein red, have a soapy touch and bitter taste They neutralize acids to form salts They contain hydrogen and oxygen as essential parts They are generally prepared by the action of a basic oxide with water

Neutralization is the process by which a salt is prepared by the reaction of an acid and an alkali

A salt is a chemical compound formed by the action of an acid on an alkali, in this way replacing the hydrogen of the acid with the metal

Salts are prepared by the neutralization of acids and alkalies, by the action of a metal on an acid, by the action of an acid on the oxide or hydroxide of a metal, by the action of an acid on the salt of another acid, by double decomposition, by the direct union of a metal and a non-metal and by the displacement of a metal from its salt by another metal

QUESTIONS

1 Explain the term 'acid' State the general properties of acids

2 What is a base and an alkali ? Give the characteristic properties of alkalies

3 Give, with examples, the general methods of the formation of acids and alkalies

4 Define the term 'neutralization' What is the result of neutralization ? Give other methods of the preparation of salts Illustrate your answer with examples

5 What is a salt ? Give the method of preparation of a salt you might have prepared yourself in the laboratory

6 What happens when ,—

(i) Silver nitrate is added to sodium chloride ?

(ii) Nitric acid is added to lead oxide (litharge) ?

(iii) Sulphuric acid is added to zinc oxide ?

(iv) Sulphuric acid on sodium sulphite

7 How can you show that —

(i) Soap contains an alkali ?

(ii) Sputum contains an alkali ?

(iii) Cow's milk has an acid ?

(iv) Urine has an acid ?

(v) Vinegar is an acid ?

What happens when vinegar is added to limestone and why ?

8 What is the difference between an oxide and a hydroxide ? What is formed when an acid acts on each ?

9 What do you understand by acid anhydrides and basic oxide? What happens when they are dissolved in water ?

10 You are given four bottles containing water, a solution of caustic soda, dilute hydrochloric acid and solution of sodium nitrate What tests will you try to distinguish one from the other ?

11 How would you classify the following compounds ?

(a) Ammonium Hydrate , (b) Lime water , (c) Copper oxide (d) Carbonic acid , (e) Hydrochloric acid , (f) Sodium chloride , (g) Silver nitrate , (h) Magnesium oxide , (i) Caustic soda , (j) Alum , (k) Carbon dioxide and (l) water

12 Given sulphuric acid and caustic soda, how will you proceed to get sodium sulphate ?

CHAPTER XV.

METALS

It has been already stated that substances exist in the form of elements and compounds. On account of certain properties, chiefly physical, common to a large number of elements, and more or less absent in others, the elements are divided into two classes, known as metals and non-metals. We have, so far, studied a brief account of a few important non-metals, such as, oxygen, nitrogen, hydrogen, sulphur, carbon, phosphorus and chlorine, etc. In this part of the book we will discuss briefly the chemistry of a few important metals and their salts.

Metals and Non-metals The division of elements into metals and non-metals, as has been stated above, is chiefly based upon the physical properties. Thus substances which possess a shining lustre, are hard and fairly heavy, can be beaten into plates and drawn into wires and are good conductors of heat and electricity, are classified as metals, and others in which these properties are absent are termed as non-metals. But the dividing line between these two groups cannot be a sharp one, as these two classes gradually merge into one another, and certain elements are sometimes placed in one class and sometimes in the other. Arsenic and antimony show some properties common to metals and non-metals. Such elements have been called metalloids, *i.e.*, elements which have properties characteristic of both metals and non-metals So we see that the division of elements into metals and non-metals

though convenient is arbitrary, because there are exceptions to most of the definitions of each. The following table gives the main features of difference, with the important exceptions :—

	Metals	*Non-metals*
Physical.—(1)	Solid at ordinary temperature except mercury which is a liquid	Non-metals are generally solids and gases. Bromine is a liquid
(2)	Have a high specific gravity but aluminium is 2 65 times heavier than water, while sodium and potassium are lighter than water	Low specific gravity
(3)	High melting point, *e g*, copper melts at 108°C iron at about 150 C but tin melts at only 23° C, while mercury is a liquid	They are chiefly gases & solids The solids have a low melting point Carbon has not been melted so far
(4)	Opaque but thin sheets of gold are translucent	Transparent (gases) and translucent. Some (solids) are opaque
(5)	Metallic lustre, *i e*, reflect light from polished or freshly cut surface	Do not possess this property Graphite and iodine are shining when freshly cut
(6)	Malleable, *i e*, can be beaten into sheets and ductile, *i e*, drawn into wires, but bismuth is brittle	Solids are brittle
(7)	Good conductors of heat and electricity	Poor conductors, except graphite Hydrogen is a good conductor of electricity
(8)	Have tenacity	Possess no tenacity

	Metals	Non-metals
Chemical,—(1)	Form basic oxides but zinc oxide and aluminium oxide act as feebly acidic oxides in presence of strong bases. Higher oxide of manganese is acidic	Form acidic oxides. But Nitrous oxide is not acidic
(2)	Generally dissolve in mineral acids giving off hydrogen	Do not usually dissolve easily in mineral acids
(3)	Either form no compounds with hydrogen or form unstable compounds which are usually non-volatile	Form stable compounds with hydrogen which are usually volatile

A study of the above table makes it clear that it is difficult to draw a hard and fast line between metals and non metals and that the division is a rough and arbitrary one. It is followed because it is convenient.

Occurrence of Metals. Some metals occur *free* in nature, or as they are said, in the *native* condition. Important among these are gold, silver, copper, lead, mercury, and platinum. Most of the metals, however, are present in the form of their compounds with other metals and non-metals in the earth's crust. These compounds are known as minerals: *e.g.*, iron pyrites, zinc blende, galena, etc. Those minerals from which metals can be easily extracted without much labour and cost are termed as *ores* of the metals, *e.g.*, iron pyrites is ore of iron, zinc blende of zinc, galena of lead, and so on. These ores are mostly oxides, sulphides, carbonates, sulphates, or silicates. The process of obtaining metals from their ores is called

metallurgy Most of the ores are converted into oxides and reduced. But the modes of extraction of metals differ with different metals

Alloys If two or more metals are melted together they usually, but not always, form a homogeneous liquid The metallic mixture got on solidification in such cases is known as an alloy The alloys are generally harder and have a lower melting point than the metals of which they are mixtures Their use depends upon the property they possess, *e g*, gold coin is made of gold and copper because the alloy is hard, type metal is used in making type letters because it expands on cooling and is easily fusible, and so on.

An alloy which contains mercury as one of the metals is known as an amalgam, as sodium amalgam, which is used for preparing hydrogen from sodium, and zinc amalgam used in electric cells to avoid local action and tin amalgam used in mirrors

The most important alloys with their important properties are given below

1. *Brass* It consists of two parts of copper to one of zinc It is more fusible and harder than copper. It casts well and is hence used in making vessels and parts of machinery.
2. *Bronze* Common bronze contains 90 parts of copper and 10 of tin, but one used for coins contains 95 parts of copper, 4 of tin and one of zinc It is easily fusible, casts well and does not corrode It is used for castings - statues, coins, utensils, etc.
3. *Dutch metal*—Contains zinc one part with copper four parts. It is very malleable and its leaves are used in electroscopes

4. *Bell metal*—Contains copper four parts with one part of tin. It is hard, brittle and sonorous Used in making bells

5. *Gun-metal*—Contains nine parts of copper with one part of tin It is very tough and tenacious and is used in guns.

6. *German-silver.* Contains copper three parts, zinc two parts, and nickel one part It is tough and malleable and white in colour. Used in coins and ornaments

7. *Type metal* -Contains four parts of lead with one of antimony It is easily fusible and expands on cooling Hence it is used in printing presses

8. *Pewter*—Contains four parts of tin with one of lead. It is easily fusible and expands on cooling

9. *Solder*—Contains equal parts of tin and lead. But the composition varies. It is soft and its melting point is very low.

10. *Gold coin*—Contains gold eleven parts with copper one part. It is harder than gold.

11. *Silver coin*—Contains eleven parts of silver.

12. *Steel*—Contains iron and some other metals *e.g.*, manganese, nickel, chromium, tungsten, etc. The proportions differ with different kinds

SUMMARY

Elements are divided for the sake of convenience into metals and non-metals

Metals are generally hard, opaque, solid at ordinary temperature, heavy, good conductors of heat and electricity and when freshly

but show a metallic lustre. They are generally malleable, ductile and tenacious. They form basic oxides, dissolve in acids with the evolution of hydrogen and either do not combine with hydrogen or form unstable compounds with it

Metalloids are those elements which show the properties of both metals and non metals

Metals occur either in the native state or in the form of their salts The minerals or rocks from which metals can be easily and economically extracted are called ores The ores are generally oxides, sulphides, carbonates and sulphates

When two are more metals are melted together they form an alloy. The alloys are generally harder and have a lower melting point than the pure metals. Amalgam is an alloy when one of the metals is mercury

QUESTIONS

1 Give the characteristic properties of metals and compare them with those of non metals

2 What are the principal characteristics of metals? Do you consider the division of elements into metals and non-metals satisfactory? *(P U M L S L C, 1923)*

3 Classify the following elements into metals or non-metals, carbon, nitrogen, copper, graphite, lead, chlorine, iodine, antimony, arsenic

4 Define the terms —

Ore, Mineral, Alloy, Amalgam, Metallurgy and Native metal

5 Name some important alloys you know Give their composition and uses.

6 Distinguish between —

(1) Alloy and amalgam.

(2) Metals and non-metals

7 State what you know about the occurrence of metals

CHAPTER XVI.

ALKALI METALS.

Sodium and Potassium.

Sodium. *Occurrence.* Sodium does not occur free in nature as it has a strong affinity for oxygen. In the combind state it is present as sodium chloride in sea water, brine springs, certain lakes and in some countries as rock salt. Sodium nitrate is another compound of sodium which occurs in large deposits in Chile as Chile saltpetre.

Preparation. Sodium is obtained by the electrolysis of fused caustic soda in cylindrical iron vessels. Caustic soda breaks up into sodium, hydrogen and oxygen. Oxygen escapes at the anode and sodium and hydrogen at the cathode. It can be prepared by the electrolysis of common sălt as well.

Properties Physical. Sodium is a silvery-white shining metal, of specific gravity 0·97. It is soft and can be cut with a knife. It melts at 96°C. and boils at 877·5°C. It dissolves in liquid ammonia giving a blue colour to the solution.

Chemical. It is a very active metal and this is why the freshly cut surface at once tarnishes—it is due to the formation of a brown film of oxide. The metal is, therefore, kept under naphtha or kerosene oil which contains no oxygen. It decomposes water at the ordinary temperature forming caustic soda and setting free hydrogen :

Sodium + water = Sodium hydroxide + hydrogen

It burns in air when heated forming sodium peroxide. It burns in chlorine to form sodium chloride.

Uses. It is used to purify ether by removing moisture from it. Sodium amalgam is used in the laboratory in the preparation of hydrogen. It is used in the manufacture of organic compounds, *e.g.* artificial rubber.

Note. Sodium should not be touched with hand because our hands are generally moist on account of perspiration.

Compounds of Sodium.

Sodium chloride is the most widely distributed salt of sodium. It has been already discussed.

Sodium hydroxide or Caustic soda. *Preparation.* (1) By the action of sodium on water. But this method is costly and not employed.

(2) In India it is prepared by boiling a solution of sodium carbonate and slaked lime. The reaction that takes place is :

$$\text{Sodium carbonate} + \underset{\text{(slaked lime)}}{\text{calcium hydroxide}}$$

$$= \underset{\text{(caustic soda)}}{\text{Sodium hydroxide}} + \underset{\text{(chalk)}}{\text{calcium carbonate}}$$

Chalk thus formed is insoluble and settles down in the vessel. The solution is filtered and caustic soda obtained by evaporation.

(3) Now-a-days when electric power is cheap, caustic soda is mainly prepared by the electrolysis of fused sodium chloride (common salt). Sodium is liberated at the cathode and chlorine at the anode. The sodium thus liberated reacts with water to form

caustic soda and hydrogen Special steps are taken to keep chlorine from coming in contact with sodium

Properties. It is a white, slightly translucent, fusible solid. It is a deliquescent substance, i.e., absorbs moisture from the air and a little carbon dioxide as well. It is soluble in water, the water becomes soapy and highly alkaline Much heat is produced when it is dissolves in water. It has a corrosive action on the skin and hence it is known as caustic soda The solution corrodes glass. Wool dissolved in it but cotton is insoluble and hence its solution is used to test the purity of woollen cloth

Uses It is chiefly used in the manufacture of soap and paper pulp.

. **Sodium Nitrate** It is found in an impure form in Chile. It contains both soluble and insoluble impurities The crude nitre is dissolved in water and filtered to separate insoluble impurities. The clear filtrate is then concentrated and sodium nitrate crystallized out. The purified sodium nitrate is then called Chile saltpetre.

Properties It is a white crystalline substance, every soluble in water. It is a hygroscopic substance and is, therefore, not used in gunpowder or fireworks When heated it gives off oxygen and sodium nitrite is left behind.

Uses It is chiefly used in the manufacture of nitric acid. It is used as a fertilizer' Much of it is converted into potassium nitrate by treating it with a concentrated solution of potassium chloride

Potassium chloride+sodium nitrate=Potassium nitrate+sodium chloride.

Potassium nitrate and sodium chloride are separated by the process of fractional crystallization and potassium nitrate is further purified.

Sodium carbonate or Washing soda. It is manufactured in two ways :—

(1)*The Le Blanc's Process.* In this process sodium chloride is first converted into sodium sulphate (salt cake) by heating it strongly with sulphuric acid

Sodium chloride + sulphuric acid = Sodium sulphate + hydrochloric acid gas.

The hydrochloric acid gas produced in the process is converted into hydrochloric acid, so hydrochloric acid is a by-product in this industry. Sodium sulphate or salt cake, as it is called, is mixed with limestone and charcoal and roasted

Sodium sulphate + calcium carbonate + carbon = Sodium carbonate + calcium sulphide + carbon dioxide.

The resulting mixture contains sodium carbonate and calcium sulphide and is known as blackash. This is then treated with warm water to dissolve out sodium carbonate which is further purified by crystallization.

(2) *Solvay or Ammonia Soda Process.* In a strong solution of common salt saturated with ammonia, carbon dioxide under pressure is passed when sodium bicarbonate and ammonium chloride are formed :

Sodium chloride + ammonia + carbon dioxide.
= Sodium bicarbonate + ammonium chloride

Sodium bicarbonate being comparatively insoluble is then separated from ammonium chloride which

remains in solution The bicarbonate is then heated and converted into sodium carbonate :

Sodium bicarbonate=Sodium carbonate+water
+carbon dioxide.

Properties. It is a white crystalline solid very soluble in water. The crystals contain a very large amount of water of crystallization. The crystals effloresce on exposure to air, *i.e.*, lose their water of crystallization and fall into an amorphous powder

Uses. It is largely used for washing purposes and in the manufacture of soap. Its crude form is known as *sajji* in the Punjab.

Sodium bicarbonate or Baking soda It contains sodium, carbon, hydrogen and oxygen as its constituents.

Preparation. (1) It is prepared as a by-product in the ammonia soda process described above.

(2) It is prepared by passing carbon dioxide in a solution of washing soda, so that reaction takes place :

Sodium carbonate+carbon dioxide
=Sodium bicarbonate.

Properties It is a white powder and is much less soluble in water than washing soda. When heated, it gives off carbon dioxide and is converted into carbonate :

Sodium bicarbonate=Sodium carbonate
+carbon dioxide+water.

It is, therefore, used in making baking powder Baking powder contains sodium bicarbonate and tartaric acid which do not react when dry but in the

presence of water they react to produce carbon dioxide the bubbles of which expand by heat on baking. Baking powder is used for cooking bread (*double roti*). It is used in medicines, as a cure for indigestion and in fire extinguishers. In Indian households it alone and not baking powder, is used to boil grains, etc.

Sodium sulphate or Glauber's Salt contains sodium, sulphur and oxygen. It is obtained by the action of sulphuric acid on sodium chloride as a by-product in the manufacture of hydrochloric acid

Sulphuric acid + sodium chloride.
= Sodium sulphate + hydrochloric acid gas.

It is a transparent, crystalline substance containing a large amount of water of crystallization which it loses on exposure to air It occurs in the United Provinces and Bihar and is named as *"Patna Khari"*

Uses It is used in the manufacture of sodium carbonate and glass. It is used in medicine as a purgative

Glass When mixture of sodium carbonate and calcium carbonate is fused with pure sand (quartz sand), the product is a transparent mass which is insoluble in water or acids and is the ordinary window glass. Instead of sodium carbonate, sodium sulphate mixed with carbon is used. The ratio generally is 100 parts of sand, 35 to 40 of soda and 15 parts of limestone. Hard glass or Potash glass is obtained by using potassium carbonate instead of soda. Special varieties of glass, *e.g.*, Flint glass, Jena optical glass and coloured varieties are prepared for various purposes

Potassium

Occurrence. Like sodium it does not occur free in nature but is present in the form of its salts. The double chloride of potassium and magnesium found in

Germany in the Stassfurt deposits are known as Carnallite. In India potassium nitrate is found in large quantities as 'kalmi' or 'shor'. Potassium salts are present in plants and animal tissues They are present dissolved in sea water.

Preparation. (1) Like sodium it is also prepared either by the electrolysis of caustic potash or potassium chloride It has no use as a metal It resembles sodium in most of its properties. It is a silvery white shining metal which tarnishes soon on exposure to air. It is also soluble in ammonia forming an indigo-blue solution. It melts at 62·3 C and boils at 667°C. It is more active than sodium and acts upon water so vigorously and so much heat is produced that the hydrogen evolved catches fire It is soft and can be cut with a knife. It is also kept under kerosene oil It is lighter than water and floats on it Some of its important compounds are mentioned below.—

Caustic potash or Potassium hydroxide is prepared in the same way as caustic soda, *i e*, by the electrolysis of potassium chloride

It is a white, deliquescent substance It is very soluble in water, the solution is highly alkaline. It is corrosive It is chiefly used in the manufacture of toilet soap

Potassium nitrate—nitre (*Shora*) This occurs as a white crust in India (*Shor*). It is obtained from the earth as follows -

The earth containing nitre and other soluble impurities chiefly common salt is dissolved in water and allowed to stand The liquid is then made to pass over thatched roofs to filter it and to remove suspended impurities The clear filtrate is led into shallow beds where the water evaporates in the sun. The concentrated liquid is then heated in large iron pans to

prepare a saturated solution. The saturated solution is then crystallized. Nitre being more soluble in hot than in cold water separates out first. It is further purified.

(2) It is obtained from Chile saltpetre as explained already.

Properties and Uses. It is a white, crystalline solid. It is highly soluble in water. The temperature of water falls considerably when it dissolves in it. When it is heated strongly it decomposes into potassium nitrite and oxygen. It is not hygroscopic like sodium nitrate and is therefore used in the manufacture of nitric acid, gunpowder, firework and matches. Gunpowder as already stated in a previous chapter is a mixture of 15 parts of nitre, 3 parts of charcoal and 2 parts of sulphur The ingredients are moistened with water and thoroughly ground. It is an explosive powder. The explosive nature is due to the sudden evolution of large volumes of carbon monoxide, carbon dioxide and nitrogen when it is ignited.

Potassium chloride. As stated above this salt is found in Germany and is obtained from carnallite. It is a white, crystalline substance, readily soluble in water. It is used as a fertilizer and in the preparation of caustic potash and other salts of potassium.

Potassium carbonate (*Khar.*) *Preparation.* (1) Formerly it was prepared from wood ashes. The ashes are treated with water and filtered. The filtrate is evaporated to dryness. The salt obtained is crude and is called potash It is purified to get pearl ash.

(2) It is now largely prepared by the Le Blanc's process explained above, in the manufacture of sodium carbonate. The Solvay process cannot be used, as potassium bicarbonate is highly soluble in water

Properties and Uses. It is a white solid. It is deliquescent and is readily soluble in water. The solution is alkaline in properties. It is used in making soft toilet soap, and in the manufacture of glass. It is used as a fertilizer in India.

Potassium chlorate. *Preparation.* It contains potassium, chlorine and oxygen. It is prepared along with potassium chloride by passing chlorine into a hot solution of caustic potash. Now-a-days it is chiefly prepared by the electrolysis of sodium chloride when sodium and chlorine are liberated. Sodium dissolves in water to form caustic soda and hydrogen. The chlorine evolved acts upon caustic soda and produces sodium chlorate and sodium chloride. The two are separated by fractional crystallization. The sodium chlorate is mixed with equal quantity of potassium chloride and potassium chlorate is obtained. Potassium chlorate can be obtained by the electrolysis of potassium chloride directly, but this is not done as there is a huge wastage of electricity.

Properties and Uses. It is a white, crystalline substance. When heated it gives out oxygen, and potassium chloride is left behind. It is used in the preparation of oxygen and manufacture of fire works and matches. It is used in medicine, especially in cases of throat trouble.

Potassium hypochlorite contains less oxygen than chlorate. It is obtained along with potassium chloride by passing chlorine into cold caustic potash solution. It sets free chlorine when acted upon by an acid.

Potassium permanganate (*Lal dawai*) contains potassium, manganese and oxygen as its constituents. It is a deep purple-red crystalline substance. When dissolved in water, it gives a deep purple solution. It is a strong oxidizing agent and is used as a disinfectant. It is used for disinfecting water, especially in cholera days.

It is prepared by fusing manganese dioxide, with potassium carbonate in the presence of air or an oxidizing agent (nitre) In this way we get potassium manganate which is then treated with ozone to get potassium permanganate. When it is heated it gives up oxygen

Potassium dichromate (*Surakh kahi*) is an orange-red, crystalline substance, soluble in water It is very powerful oxidizing agent. When heated, it decomposes giving off oxygen and a green chromium salt is left behind It is used in tanning leather and as a pigment.

Other important salts of potassium are potassium iodide and potassium bromide used in medicine

SUMMARY

Potassium and sodium are known as alkali metals They do not occur free in nature but their salts are found in the earth's crust or in sea water The two metals and their compounds are similar to each other The metals are obtained by the electrolysis of their chlorides or hydroxides

Sodium and potassium are silvery white, shining metals which tarnish soon on exposure to air They are active, light, soft and decompose water even at the ordinary temperature Therefore, they are kept under kerosene oil Sodium when thrown in water, liberates hydrogen and forms caustic soda In the case of potassium the action is so vigorous and so much heat is produced that hydrogen liberated catches fire The metals are important on account of their compounds

The important compounds of sodium are

(1) Sodium chloride which contains sodium and chlorine

(2) Sodium hydroxide which contains sodium, hydrogen and oxygen

(3) Sodium carbonate which contains sodium, carbon and oxygen

(4) Sodium bicarbonate which contains sodium, carbon, hydrogen and oxygen

(5) Sodium nitrate which contains sodium, nitrogen and oxygen

The important salts of potassium are potassium chloride potassium hydroxide, potassium carbonate, potassium nitrate, potassium chlorate (potassium, chlorine, oxygen) and potassium dichromate (potassium, chromium, oxygen)

Glass is obtained by fusing together sodium carbonate or potassium carbonate, lime stone and sand There are varieties of glass prepared for different purposes.

QUESTIONS

1 Write what you know about the occurrence, preparation, properties and uses of the metal sodium

2 How was caustic soda prepared in old days? How is it prepared now? Give its properties and uses

3 How is sodium carbonate prepared from common salt? Give its properties and uses

4 Define the term by product Name any substance which is produced as a by-product in some industry

5 What is the composition of baking powder? On what does its use in baking depend?

6 How do sodium and potassium resemble each other? Give the important properties of potassium

7 Write what you know about the composition, properties and uses of the following (a) caustic soda, (b) Glauber's salt, (c) brass
(P U, M & S L C, 1925)

8 Describe a cheap method of preparing caustic soda in large quantities and explain the chemical changes which occur in this process State the important properties and uses of the substance Plaster of Paris (U P, M & S L C 1918)

9 How is nitre met with in nature? How is it purified? To what use is it put? Give its properties

10 Give the composition of gun-powder On what does the explosive power of the powder depend?

11 What is glass? How is it prepared?

12 Give the composition, preparation, properties and uses of the following salts (a) caustic potash, (b) potassium carbonate, (c) potassium chlorate, (d) potassium permanganate, (e) potassium dichromate

13. What happens when—

(a) solution of sodium nitrate and potassium chloride are treated together

(b) potassium chlorate is heated

(c) potassium permanganate is heated

(d) sodium nitrate is heated?

14 Write a short account of the properties and uses of the following

(a) potassium permanganate, (b) saltpetre, (c) nitric acid,
(d) bleaching powder (P U, M & S L C, 1928)

15 What happens when (a) hydrochloric acid gas, and (b) carbon dioxide is passed through a solution of caustic soda?

16 Starting with sodium chloride how would you prepare (a) sodium carbonate, (b) caustic soda, (c) chlorine?

CHAPTER XVII.

COPPER, SILVER AND GOLD

Copper (*Tamba*) *Occurrence.* Copper occurs free in nature and was known to the ancients. In the native condition it is found in America, Japan, Australia and India. Its chief ores are —

(1) Copper pyrites—contains copper, iron, sulphur

(2) Cuprite or ruby copper,—contains copper and oxygen,

(3) Copper glance,—contains copper and sulphur.

(4) Malachite,—contains copper, carbon and oxygen.

Properties. Physical. It has a characteristic red colour. Its specific gravity is 8·95 and melting point is 1083° C. It is very malleable and ductile, and hence can be rolled into sheets, hammered into leaves and drawn into wires. Next to silver, it is the best conductor of heat and electricity.

Chemical In pure, dry air copper does not tarnish (oxidize), but in ordinary moist air the metal rapidly tarnishes and becomes covered with a brown film of oxide or sulphide. On prolonged exposure, it becomes coated with a green film called verdigris (carbonate of copper). This layer protects the rest of the metal under it from being corroded. On heating in air, the reddish surface of copper becomes covered with black scales of copper oxide.

Hydrochloric acid, dilute and cold sulphuric acid do not act upon the metal. Strong, hot hydrochloric acid in the presence of air acts upon it very slowly to produce cupric chloride; in the absence of air no action takes place:

Hydrochloric acid+copper+oxygen of air
=cupric chloride+water.

Hot concentrated sulphuric acid readily dissolves the metal forming copper sulphate and sulphur dioxide:

Copper+sulphuric acid=Copper sulphate+sulphur dioxide+water.

Nitric acid, dilute or strong, acts upon it at once even when cold, forming copper nitrate and red oxide of nitrogen:

Copper+nitric acid=Copper nitrate+nitrogen peroxide+water.

Uses. Next to iron it is the most widely used metal. On account of its cheapness and great conductivity of electricity it is largely used in electric wires and cables. It is used in making utensils, coins, boilers, stills. It is used for sheathing ships and for bolts, because it resists corrosion by sea-water.

Alloys The important alloys of copper with other metals are:—

(1) Brass (*pital*) contains copper 2 parts with zinc 1 part.

It is used in making utensils, idols and in machinery.

(2) Bronze (*kansi*) contains 9 parts of copper and 1 part of tin. It is also used in making household utensils, statues, etc.

(3) German silver contains copper, zinc and nickel, and is used for making utensils and coins.

Salts. The most important salt of copper is copper sulphate (blue vitriol—*nila totia*) It is prepared by heating copper with strong sulphuric acid. It is a blue crystalline substance containing a large amount of water of crystallization. When heated, the water of crystallization is given off and a white anhydrous powder is left behind which again becomes blue when water is added to it.

EXPERIMENT 131. Powder, a few crystals of copper sulphate and heat the powder in a test tube. Drops of water condense on the upper parts of the tube. The powder becomes whitish in colour. Cool the tube and add a few drops of water. The powder again becomes blue.

Copper sulphate is used in batteries (Daniell Cell) and in copper-plating. It is used in the preparation of nearly all other compounds of copper. It is used in calico printing, in the preparation of certain pigments. It is used in medicine as a caustic in granular lids It is mixed with lime (both in solution form) for spraying plants to prevent the growth of fungi. It is used to destroy weeds and algæ in standing water.

Silver (*Chandi*)

Occurrence. It has been known from the earliest times as it occurs free in nature. It is generally present, mixed with small quantities of gold and copper It is present in lead ores. Its important ores are:—

(1) Silver glance, which contains silver and sulphur.

( ) Horn-silver, which contains silver and chlorine

(3) Galena, an ore of lead, contains silver from which it is chiefly extracted.

Properties. Silver is a white metal with specific gravity 10·5. It is very malleable and ductile. It melts at 960°C. It is the best conductor of heat and electricity.

It is not oxidized by pure air but in air which contains traces of sulphuretted hydrogen it becomes black on account of the formation of black sulphide

Hydrochloric acid has no action on silver; strong and cold sulphuric acid has no action, but when heated it dissolves silver forming silver sulphate and sulphur dioxide:

Sulphuric acid + Silver = Silver sulphate + sulphur dioxide.

Nitric acid dissolves it readily forming silver nitrate and oxide of nitrogen:

Silver + nitric acid = Silver nitrate + oxide of nitrogen + water

Uses It is used in making ornaments, utensils, mirrors, idols and for other decoration purposes. Alloyed with copper it is used in making coins and ornaments. It is used in electroplating base metals.

Salts. Silver Nitrate is the most important salt of silver and is obtained by the action of nitric acid on silver. It is a colourless, crystalline substance and is readily soluble in water. The solution when exposed to sunlight becomes black. It is, therefore, kept in dark bottles. When a drop of the solution

falls on cloth or skin, a black stain is produced. It is, therefore, used in marking-inks and hair-dyes. Its solution is used in medicine as caustic lotion. It is used as a reagent in the laboratory as a test for chlorides and hydrochloric acid.

Other salts of silver, *e.g.*, silver chloride, silver bromide, silver iodide also turn black on exposure to light and are, therefore, used in photography.

Gold (*Sona*). *Occurrence.* It is one of the first metals known to man. It is found usually in the native condition mixed with a little amount of silver and copper. It is worked in India in Mysore.

Properties. It has a beautiful yellow colour, is a brilliant metal It is the most malleable and ductile of all the metals. Its specific gravity is 19·43. It melts at 1075°.

It is not affected by air or by sulphuretted hydrogen. It is not soluble in any single acid but dissolves in *aqua regia.* It combines directly with chlorine and bromine.

It is too soft for use as ornaments or for coinage and is, therefore, alloyed with copper or silver or both. The alloy is harder and is able to resist wear. It is used for electroplating silver and other metals.

SUMMARY.

Copper, silver and gold occur free in nature and were known to ancients All these metals are found in India The important ores of copper are copper pyrites, copper glance, cuprite and malachite The important ores of silver are silver glance and horn-silver

Silver is the best conductor of heat and electricity, and next comes copper Both are very ductile and malleable Both have a high melting point The action of hydrochloric acid, sulphuric acid and nitric acid is the same on the two metals

Gold is not soluble in any single acid but dissolves in aqua regia It remains unaffected in air

The important salt of copper is copper sulphate and that of silver is silver nitrate Silver salts turn black on exposure to light and are, therefore, used in photography

QUESTIONS

1 What do you know about the occurrence of copper ? Name some of its ores. Describe its physical properties

2 What is the action of important acids on copper ? What is the effect of air on copper ?

3 Describe the preparation, properties and uses of copper sulphate,

4 How does silver occur in nature ? Describe its properties and uses

5 Describe the preparation, properties and uses of silver nitrate Why are silver salts used in photography ?

6 Describe the properties and uses of gold

7. How does copper occur in nature ? Describe the properties and uses of copper and copper oxide Mention some of the more important alloys of copper What happens when nitric acid is added to copper ? *(P.U , M, & S L,C , 1913)*

8 Explain what happens when copper pice and silver rupee are exposed to air for a sufficiently long time ? Why does a silver spoon become tarnished when placed in beaten egg ?

9 What happens when a crystal of copper sulphate is heated, and when water is poured over the heated crystal ?

10 How do copper and silver resemble each other ?

CHAPTER XVIII.

ZINC MERCURY, MAGNESIUM

Zinc (*Jast*). *Occurrence.* Zinc does not occur free in nature. The important ores of zinc are :—

(1) Zinc blende (zinc sulphide), contains zinc and sulphur.

(2) Zincite (zinc oxide), contains zinc and oxygen and

(3) Calamine (zinc carbonate), contains zinc, carbon and oxygen.

Properties. Zinc is a bluish white metal. At ordinary temperature it is brittle but becomes soft, malleable and ductile when heated. It melts at 419°C. Its density is 7.1. It is not easily attacked by air and hence is used for coating iron sheets to protect them from rusting. Iron sheets are cleaned and dipped in molten zinc, when a thin layer of zinc adheres to the iron. The porcess is known as galvanizing iron and the iron is known as *galvanized* iron.

When heated strongly zinc burns in air forming a white powder—zinc oxide. Zinc dissolves in dilute and concentrated hydrochloric acid to form zinc chloride and hydrogen. Dilute sulphuric acid acts on it to form zinc sulphate and hydrogen but hot concentrated sulphuric acid produces sulphur dioxide. It is readily soluble in nitric acid.

Zinc + dil hydrochloric acid = Zinc chloride + hydrogen.

Zinc + dil. sulphuric acid = Zinc sulphate + hydrogen
Zinc + con. sulphuric acid = Zinc sulphate + sulphur dioxide + water.

Zinc + nitric acid = Zinc nitrate + oxides of nitrogen + water.

Pure zinc is not acted upon by hydrochloric acid Zinc is soluble in caustic potash and caustic soda and evolves hydrogen from them (*See* Hydrogen)

Uses Zinc is used in galvanizing iron. It is used as a negative pole in batteries Its important alloys are Brass, Dutch metal and German silver.

Salts. *Zinc oxide* It is prepared by burning zinc in clay crucibles (*Kuthali*) It is powdered and used in medicine, paints and pigments

(2) *Zinc sulphate.* It is prepared by the action of zinc on sulphuric acid It is a cryatalline substance and contains a large amount of water of crystallization. It is known as white vitriol. Its solution is used as eye lotion as a cure for sore eye

Mercury. *Occurrence.* It occurs in small quantities in the free state and was known to the ancients under the name of quicksilver as its colour is silvery white and it flows like water It cannot be picked up by hand if thrown on the ground or any other surface.

Its important ore is Cinnabar = Mercury sulphide (mercury and sulphur).

Properties It is the only metal, which is liquid at the ordinary temperature. It is silver white in colour with specific gravity 13·59 It solidifies at 39° C and boils at 357°C. Impure mercury leaves a tail behind when rolled over a smooth surface of glass

The metal is not affected by ordinary air, but when heated it combines with oxygen, forming red oxide of mercury. The red oxide of mercury if heated decomposes into mercury and oxygen.

Mercury is not acted upon by dilute or strong hydrochloric acid and dilute sulphuric acid. Hot concentrated sulphuric acid acts upon it to form mercury sulphate and sulphur dioxide. Nitric acid dissolves it to form mercury nitrate and oxide of nitrogen :

Sulphuric acid + mercury = Mercury sulphate + sulphur dioxide + water.

Nitric acid + mercury = Mercury nitrate + oxide of nitrogen + water.

Purification. Commercial mercury usually contains lead, zinc and copper as impurities. It may be cleaned by filtration through chamois leather and then purified by running it several times in a thin stream through a long column of dilute nitric acid containing a little mercurous nitrate as shown in the figure (65). It is then washed with water and dried. To get chemically pure mercury, it is distilled in vacuum.

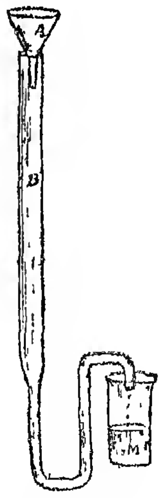


Fig. 65.

Uses. It is extensively used in barometers and thermometer. It dissolves many metals, *e.g.*, zinc, sodium, copper, etc., and is used to form their amalgams. Tin amalgam is used in mirrors. Amalgams of gold, copper and zinc, are used in dentistry. It is used in the

laboratory and in the separation of gold and silver from their ores.

Compounds : Calomel or Mercurous chloride. It contains mercury and chlorine. It is a white powder and is used in medicine in small doses as a purgative.

Mercuric chloride or Corrosive sublimate It contains more chlorine than calomel It is a poisonous substance and its weak solution is used as an antiseptic. It is sold in the form of blue pills. It eats away metals and, therefore, its solution is prepared in chinaware or glass vessels. Its antidote is egg albumen

Magnesium *Occurrence.* It does not occur free in nature, but is widely distributed in the form of magnesite (magnesium carbonate); dolomite (magnesium and calcium carbonate). In the Stassfurt deposits magnesium sulphate and magnesium chloride are also found. Talc, soap-stone and asbestos also contain magnesium. It is present in the sea-water as magnesium sulphate and magnesium chloride. Certain mineral springs contain Epsom salts (magnesium sulphate) dissolved in them.

Preparation. It is prepared by the electrolysis of fused magnesium chloride and is pressed into ribbon

Properties It is a silvery white metal, with specific gravity of 1·75. It melts at 651 °C. It is unaffected by dry air but in ordinary moist air it soon becomes tarnished due to the formation of a film of the oxide and basic carbonate of magnesium. It burns in air with a dazzling white light. It burns in nitrogen also forming magnesium nitride.

The metal dissolves in dilute acids but not in alkalies. It is the only metal which evolves hydrogen from dilute nitric acid.

Uses. Magnesium powder mixed with powdered potassium chlorate burns with a noise when lighted, producing a blind white flash, used in photography. The metal is a constituent of some light alloys used in motor pistons.

Compounds. Magnesium sulphate or Epsom Salt is the most important salt of this metal It occurs in sea-water and in certain springs It is prepared by the action of magnesium on sulphuric acid It is a white, crystalline substance containing a large amount of water of crystallization It effloresces when exposed to air. It is used in medicine as a purgative under the name of magnesia and is also used in dyes.

SUMMARY

Zinc and magnesium do not occur free in nature The ores of zinc are zinc blende, calamine and zincite Magnesium is widely distributed in nature as magnesium carbonate and sulphate It is present in sea water in the form of magnesium sulphate and chloride

Mercury occurs in free State Its ore is cinnabar

Zinc and magnesium burn in air forming their oxides Mercury when heated combines with oxygen to form red oxide of mercury which on further heating decomposes into oxygen and mercury

Mercury is purified by treating with dilute nitric acid It is the only liquid metal

The important salts of zinc are zinc sulphate and zinc oxide Both are used in medicine The salts of mercury are mercurous and mercuric chlorides The salt of magnesium is magnesium sulphate used in medicine as a purgative

QUESTIONS

1 Name the principal ores of zinc Give their composition

2 What are the chief physical and chemical porperties of zinc ? Give its uses.

3 How is galvanized iron prepared ? What is its use ?

4 What do you know about the occurrence of mercury ? Give its chief properties and uses

5. Name the important salts of zinc and mercury. Give their uses.

6. Enumerate the properties and uses of magnesium.

7. What is Epsom salt? How is it prepared? What are its uses?

8. What happens when—
   (a) zinc is heated strongly in air,
   (b) mercury is heated in air,
   (c) magnesium is heated in air,
   (d) red oxide of mercury is heated?

9. What is the action of important acids on—
   (a) magnesium,
   (b) mercury,
   (c) zinc?

10. How is mercury purified? Why is it used in thermometers and barometers?

CHAPTER XIX

Tin and lead

Tin *Occurrence.* Tin occurs in small quantities in the native condition. Its important ore is tin-stone (tin and oxygen)

Properties. Physical. Tin is a white metal with a specific gravity of 7 3. It melts at 232°C. At the ordinary temperature it is malleable but when heated to 200°C, becomes brittle. Tin foil produces a peculiar sound when bent.

Chemical Tin is not affected either by air or water. If it is heated strongly it burns with a dazzling white light forming tin oxide.

Dilute hydrochloric acid has a slow action on tin but hot and concentrated hydrochloric acid readily dissolves it, evolving hydrogen :

Hydrochloric acid + Tin = Tin chloride + hydrogen.

Dilute sulphuric acid has no action but hot concentrated acid readily attacks tin to form tin sulphate and sulphur dioxide :

Tin + sulphuric acid = Tin sulphate + sulphur dioxide + water.

Dilute nitric acid acts slowly upon tin forming tin nitrate, ammonium nitrate and water Strong acid with only a little water dissolves tin to form metastannic acid, oxide of nitrogen and water. Chemically pure nitric acid has no action on tin.

Uses As it does not tarnish in ordinary or moist air, so it is used in tinning iron sheet to protect it from rusting It is made by dipping iron sheet in molten tin. But if iron becomes exposed in a tin sheet, it readily rusts It is used in many alloys, *e g*, bronze, solder, pewter, and Britannia metal (tin, copper and antimony). It is used in tinning kitchen utensils. Tin amalgam is used in looking glasses Tin-foil is used for packing foodstuffs and other articles.

Salts. Tin oxide is used in enamels and glazes.

How to tin a vessel. Thoroughly clean and heat strongly the vessel to be tinned Rub powdered ammonium chloride on the surface with a coarse cloth, it further cleanses the surface Sprinkle a little resin on the surface, it melts and protects the surface from oxidation Now rub a piece of tin on the surface and spread it thoroughly with the piece of cloth.

Lead *Occurrence.* Lead is found free in nature in traces The chief ore of lead is Galena, the sulphide, which is a heavy mineral with a bright lustre. Galena is found abundantly in Burma Galena sometimes contains silver which is extracted by special methods.

Properties Perfectly pure lead has a bluish grey colour and possesses a silver-white lustre. It is very soft and can be cut with a knife. It marks the paper and hence was mistaken for graphite in lead pencils. It is a heavy metal with specific gravity of 11 35 and melts at 327°C. It cannot be hammered into foils or drawn into fine wires, but it can be rolled into sheets.

Chemical Freshly cut surface soon tarnishes due to the formation of a thin layer of hydroxide and carbonate. The metal is not attacked by pure water (except at the boiling point) or by air, but is rapidly

corroded in moist air or in soft water having carbon dioxide dissolved in it The hydroxide and a little carbonate formed dissolve in water and render it poisonous. Hard water containing calcium bicarbonate has much less action on lead than soft water An insoluble layer is formed in this case which protects the metal. Lead pipes, therefore, should be used for hard water and not for soft water.

When heated in air, lead first melts into a shining liquid whose surface soon becomes tarnished. Gradually it is converted into a yellow oxide called litharge (*murda zang*) On further heating this oxide takes up more oxygen and is converted into red oxide of lead known as red lead (*sandhur*). When red lead is heated, oxygen is given off and litharge is left behind. When red lead mixed with powdered charcoal is heated strongly in a clay crucible or under a blow-pipe flame, it is reduced into metallic lead

EXPERIMENT 132. Bore a hole in a piece of charcoal. Put in the cavity a little of the mixture of red lead and powdered charcoal Heat it in the reducing flame of the blow-pipe A shining bead of lead is obtained. It marks the paper and is very soft.

Hydrochloric acid, dilute or strong, and sulphuric acid, dilute, have no action on it. Strong and hot sulphuric acid dissolves lead forming lead sulphate and sulphur dioxide .

Sulphuric acid + lead = Lead sulphate + sulphur dioxide + water.

It is readily soluble in nitric acid forming lead nitrate, oxide of nitrogen and water :

Nitric acid + lead = Lead nitrate + oxide of nitrogen + water

Acids present in vegetables form poisonous compounds with lead, hence vessels which contain lead should not be used for cooking purposes. Zinc displaces lead from its salts.

EXPERIMENT 133. In a cylinder dissolve some lead acetate. Suspend in it a piece of zinc by means of a thread as shown in the diagram. Keep it for a day. A very beautiful lead tree will be obtained. Zinc displaces lead which is deposited finely on zinc. Zinc dissolves and forms zinc acetate :

Zinc + lead acetate
= Zinc acetate + lead

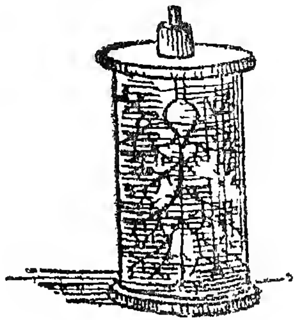


Fig 66.

Uses. It is largely used in making sheets for lining tanks and sinks and in making pipes. Lead sheet is used in lining the chambers in the manufacture of sulphuric acid. Most of the metal is used in making its salts, chiefly white lead (lead carbonate = *safeda*) used for making paints. Litharge and red lead are also used in making paints.

Alloys Solder and type metal are the important alloys. When alloyed with a little arsenic, it becomes very hard and is used in lead shot and bullets.

Compounds. The important compounds are litharge and red lead—which are its oxides.

Lead carbonate is important for its extensive use in paints. It is an amorphous powder and mixes readily with linseed oil. All the compounds of lead are poisonous.

SUMMARY

ı important ıre of tin is tinstone a d of lead, g ılena,

ıh tin and le d have a high density and low melting point ıre malleal le but not ductile Tin is unaffected by air or water s therefore used for tinning iron sheets and kitchen utensils urns in air if heated strongly to form tin oxide Tin is used in oy chiefly bronze Tin amalg m is used in making mirrors he action of acids see the table in the end]

ad when heated forms litharge and red lead Lead carbonate ese oxides are used in paints It is used in sheets and pipes by er It is used in its alloys, solder and type metals Alloyed rsenic it is used in making lead shot and bullets

the compounds of lead are poisonous

QUESTIONS

Describe the properties and uses of (1) tin, and (2) lead

Describe the action of water on lead

Write the properties and uses of (1) galena, (2) red lead arge, (4) white lead.

How does zinc occur in nature ? How would you distinguish om tin and lead ? Name some of the alloys of zinc and give r uses (P U, M & S L C, 1912)

How will you get the following —

) litharge from lead, (2) lead from red lead, (3) red lead from litharge, (4) lead from lead acetate ?

What chemical action takes place when (a) nitric acid acts b sulphuric acid on lead ?

Write short notes on—
lack lead, white lead, and red lead

CHAPTER XX.

ALUMINIUM.

Occurrence. Aluminium does not occur free in nature, but is widely distributed as a constituent of clay in the earth's crust. The most important forms are the oxide alumina and the silicates.

Corrundum, ruby, sapphire and bauxite are the chief oxides. It is profitably extracted from bauxite Ruby, sapphire, emerald, topaz and precious stones are its oxides with different colours.

Properties. Aluminium is a bluish-white metal having a specific gravity of 2·7 It is very malleable, ductile and tenacious. It is as good a conductor of electricity as silver. It melts at 658°C.

It resists the action of air its surface being covered by a thin transparent layer of oxide which prevents further attack. It does not decompose water even at the boiling point. It is acted upon by sea-water, or saltish water. Dilute hydrochloric acid acts slowly on it but hot concentrated acid dissolves it readily forming aluminium chloride and hydrogen :

Hydrochloric acid+aluminium=Aluminium chloride+hydrogen.

Dilute sulphuric acid dissolves the metal to form hydrogen and aluminium sulphate but hot strong acid forms sulphur dioxide :

Aluminium+dil. sulphuric acid=Aluminium sulphate+hydrogen

Aluminium+strong hot sulphuric acid=Aluminium sulphate+water+sulphur dioxide.

Nitric acid dilute or strong has no action on it. It is dissolved in solution of caustic soda or caustic potash with the evolution of hydrogen:

Aluminium + caustic soda = Sodium aluminate + hydrogen.

Aluminium powder is a powerful reducing agent. Oxides of certain metals, *e.g.*, copper, which cannot be reduced by charcoal are reduced at once by heating with aluminium powder.

Alloys Magnalium is an alloy with 2 to 10% of magnesium, it is lighter than aluminium and is used in making air-ships, pistons of motor engines and cheap balances.

Duralium contains aluminium, copper and magnesium

Aluminium bronze contains copper from 3 to 10%.

The alloys are hard and have a high tensile strength and are not corroded by sea-water specially if manganese is also present

Uses. Aluminium and its alloys are used for making cooking utensils, photographic cameras, telephone receivers, and picture-frames, trays, combs, and other ornamental things. When its powder is mixed with linseed oil, it is used as a white paint. Aluminium powder is used as a reducing agent in the metallurgy of certain metals, *e.g.*, manganese. Air ships and pistons of motor engines are made of its alloys. The powder mixed with black oxide of copper is used in making *phul jharris*.

Compounds. Natural oxides in the form of ruby emerald, sapphire are used as precious stones.

Alum (*Phatkari*), is a double sulphate of potassium and aluminium. It is a white crystalline substance with a large amount of water of crystallization. It is greatly soluble in water. When heated it gives out its water of crystallization and forms a white porous mass (*phul phatkari*). It is used as a mordant, *i.e.*, as an agent for fixing colours on cotton and for cleaning water. It is used in the tanning indusrry.

SUMMARY

Aluminium does not occur free in nature Its important ore is bauxite Ruby, emerald, sapphire are oxides and used as precious stones

It is a light metal with a high tensile strength and is not affected by air It is not attacked by nitric acid but dissolves in hydrochloric acid and sulphuric acid and in alkalies It is as good a conductor of electricity as silver It is used chiefly on account of its lightness, cheapness and high tensile strength in making air-ships, motor pistons, utensils, cameras, and other household and ornamental things Its powder is a reducing agent and is used as such

Its chief alloys are magnalium, duralumin and aluminium bronze.

QUESTIONS.

1 Describe what you know about the occurrence of aluminium

2 Give the properties and uses of aluminium.

3 Describe the action of air, sea water and important acids on aluminium

4 Why is aluminium so extensively used ? Give its uses

5 What happens when —

(1) Aluminium powder and copper oxide are heated together ?

(2) Aluminium is put in caustic soda ?

(3) Zinc is put in potash ?

6 Name the important alloys of aluminium Give their properties

7 Give the composition of alum What happens when it is heated in a test tube ?

CHAPTER XXI

IRON.

Occurrence. Iron is the most important and useful of the metals, so much so that we are said to be living in the Iron Age.

It occurs very rarely in the free state in nature. Iron is present in the soil and in the human body in the red colouring matter of the blood. The important ores of iron are :

(1) Magnetite or magnetic ore—oxide ; contains iron and oxygen. It is magnetic in properties and is known as Lodestone.

(2) Haematite—oxide of iron

(3) Iron pyrites—sulphide of iron.

(4) Clay iron stone—carbonate of iron.

Ancient Hindus were quite familiar with the extraction of this metal from its ores. The famous non rusting iron pillar at Delhi is a clear proof of their skill. Now-a-days cast-iron is prepared at Tatanagar in the Tata Iron Works.

Varieties of Iron. There are three varieties of iron. They differ from one another in properties and the percentage of carbon and other impurities. The three kinds are : (1) Cast-iron, (2) Wrought-iron and (3) Steel.

Cast-iron (Pig iron) It contains from 2·2 to 4·5 per cent or more of carbon mostly in the uncombined state and is called graphitic carbon. Besides carbon the other impurities met with are sulphur, phosphorus,

manganese and silicon. If this iron is dissolve hydrochloric acid, black particles of carbon are floating in the solution left in the bottle

Cast iron is hard, brittle and coarsely crystal Its melting point is 1100°C. Most of it is converted wrought-iron or steel It is used in casting var articles, *e.g.*, stoves, pipes, parts of machinery castings for ornaments for it expands on cooling water It cannot be used for making articles w have to stand great strain, *e.g.*, Railway lines,

Wrought-iron It is the purest variety cont ing from 0.12 to 0.25 per cent of carbon. It m at a higher temperature than cast iron (1400-150 The total impurities (carbon, sulphur, silicon phosphorus) are less than 0.5%. It is tough malleable and is sometimes called as malleable ir It can be rolled into sheets and drawn into wires can be welded, *i.e.*, two pieces can be united w red-hot by hammering them together Its pro of welding is exceedingly useful and is used by bl smiths in various ways It is used to make s articles as can be made by beating with hammer, horse-shoes, nails, spikes, etc

Steel Steel contains from 0.15 to 1.5 per c of carbon in the combind state. It is thus purer t cast-iron but impurer than wrought-iron. It cont only a little of silicon and no sulphur or phospho It may contain manganese, tungsten, chromium, ni etc., which are purposely added to produce alloy ste used for special purposes

The properties of steel depend on the percent of carbon; low carbon steels are soft, tough and duc like wrought iron – mild steel; increase in car makes it harder and brittle. If steel is heated

redness and then cooled suddenly by plunging into water, it becomes as hard and brittle as glass. The process is known as Hardening. By heating this hardened steel from 230°C. to 316°C. and allowing it to cool slowly in air, the steel acquires different properties The process is known as Annealing. Steel is used for making razor-blades, pen-knives and axes, cutlery, watch springs and swords, chisels, saws, etc. Steel is used in making artificial magnets because it retains magnetism.

Properties of iron. Chemically pure iron is not known in commerce. It can be prepared by the reduction of pure iron oxide by hydrogen or by the electrolysis of a solution of ferrous chloride. Pure iron is a soft, white metal with a specific gravity of about 8 and melting point 1510°C. Pure iron does not rust in dry air but in moist air it rusts. To protect it from rusting it is tinned or galvanized. It is the most magnetic of all the substances

Iron burns in oxygen to form iron oxide. It decomposes steam It is readily soluble in dilute or concentrated hydrochloric acid forming iron chloride and hydrogen. With dilute sulphuric acid it gives hydrogen and iron sulphate but with strong and hot acid it gives sulphur dioxide, iron sulphate and water. Nitric acid dissolves it to form iron nitrate and oxide of nitrogen and water.

Rusting of iron. In the presence of moist air iron is covered with a brown film known as rust.

This is iron oxide. Thus iron is corroded if left exposed to moist air for a long time. Dry air or water alone has no action on it. This can be proved by putting clean nails in a test tube containing boiled water and tightly corked not to allow air to pass and

placing iron nails in dry air. To protect it from rusting it is painted, varnished, enamelled, tinned or galvanized.

Salts. The important salt of iron is iron sulphate (green vitriol). It is prepared by the action of dilute sulphuric acid on iron. The crystals contain a large amount of water of crystallization. It is used in the manufacture of inks and paints and as a reagent in the laboratory

SUMMARY

There are three varieties of iron—cast or pig iron, wrought or malleable iron and steel Cast-iron contains from 3 to 4% of carbon besides sulphur, phosphorus, manganese and silicon It is brittle but casts well and is used in articles which are prepared by casting Wrought iron is the purest variety containing only 12 to 25% of carbon and other impurities It can be rolled into sheets and drawn into wires It can be welded It is used for making nails, rods, horse shoes, etc Steel contains from 15 to 1 5% of carbon without sulphur or phosphorus Its properties vary between cast iron and wrought iron Its properties depend upon the percentage of carbon and the heat treatment It is used in cutlery, razors, chisels, saws, etc Pure iron is obtained by reducing iron oxide Iron rusts in damp air It dissolves in all acids

QUESTIONS

1, Name the various varieties of iron Give their distinctive properties and important uses

2 Mention the distinguishing characteristics and the important uses of the different commercial varieties of iron What is the effect of long exposure of iron to moisture ? What devices are used to protect it from the action of moist air ? (*P U , M & S L C* , 1931)

3 What do you know about the composition, properties and uses of the following —(*a*) washing soda (*b*) tinplate, (*c*) lime and (*d*) steel (*P U , M & S L C* , 1923)

4 Enumerate the chief properties and uses of any two of the following metals —Copper, [illegible], iron and lead (*P U , M & S L C* , 1926)

5 What do you understand by the terms hardening and annealing as applied to steel ? How do they modify its properties ?

Describe the action of important acids, both strong and dilute, on iron

CHAPTER XXII.

ACTION OF IMPORTANT ACIDS, AIR AND WATER ON METALS

Action of acids We have studied in previous chapters the action of important acids on the various metals. We have seen that the action depends upon the concentration of the acids as well as its temperature In this chapter we will summarize what we have read before

Hydrochloric acid Copper, silver, mercury and lead are not acted upon by dilute hydrochloric acid, aluminium and tin are slowly attacked by the dilute acid; sodium, potassium, zinc and iron are soluble in dilute acid with the formation of their chlorides and evolution of hydrogen. Strong acid acts upon all the above metals except silver, mercury and lead giving the same products as dilute acid Copper is acted upon by concentrated acid in the presence of air.

Sulphuric acid. Dilute sulphuric acid acts on sodium, potassium, commercial zinc, iron and aluminium to form their sulphates and hydrogen Pure zinc is not attacked by either dilute hydrochloric or dilute sulphuric acid Strong and hot sulphuric acid acts upon all the metals (except gold) dealt with in this book forming sulphur dioxide, water and their corresponding sulphates

Nitric acid Nitric acid acts upon all the metals, except aluminium to form their corresponding nitrates

oxides of nitrogen and water. In the case of tin, however, concentrated acid with a little water forms metastannic acid, oxides of nitrogen and water. Absolutely pure nitric acid has no action on tin while dilute nitric acid slowly dissolves tin, forming tin nitrate, ammonium nitrate and water. Magnesium is the only metal which evolves hydrogen when acted upon by dilute nitric acid.

The following table gives the action of the three important acids on the metals studied in this book :

Metal	*Hydrochloric acid*	*Sulphuric acid*	*Nitric acid*
Copper	No action with dilute acid. It dissolves slowly in the strong acid when exposed to air and boiled. Copper + hydrochloric acid = Cuprous chloride + hydrogen	No action in the cold but hot and strong acid dissolves it giving up sulphur dioxide. Copper + sulphuric acid = Copper sulphate + sulphur dioxide + water	Reacts readily even with cold and dilute acid. Nitric acid + copper = Copper nitrate + oxide of nitrogen + water
Zinc	No action on pure zinc but impure zinc is acted upon readily by dilute or strong acid. Zinc + hydrochloric acid = Zinc chloride + hydrogen	Dilute acid gives hydrogen. Zinc + sulphuric acid (dil.) = Zinc sulphate + hydrogen. Strong and hot acid gives sulphur dioxide. Zinc + sulphuric acid (hot, strong) = Zinc sulphate + sulphur dioxide + water	Zinc + nitric acid = Zinc nitrate + oxide of nitrogen + water

Metal	*Hydrochloric acid*	*Sulphuric acid*	*Nitric acid*
Silver	No action	Cold and dilute acid has no action, but hot concentrated acid dissolves it Silver + sulphuric acid (conc)=Silver sulphate + sulphur dioxide + water	Acts readily even in the cold Silver + nitric acid = Silver nitrate + oxide of nitrogen + water
Mercury	No action	Cold and and dilute acid has no action but hot and strong acid readily dissolves it Mercury + sulphuric acid = Mercury sulphate + sulphur dioxide + water	Acts readily Mercury + nitric acid = Mercury nitrate + oxide of nitrogen + water
Sodium or Potassium	Sodium + hydrochloric acid = Sodium chloride + hydrogen	Sodium + sulphuric acid = Sodium sulphate + hydrogen	Sodium + nitric acid = Sodium nitrate + oxide of nitrogen + water
Aluminium	Dissolves slowly in dilute and cold acid but at once in strong acid . Aluminium + hydrochloric acid = Aluminium chloride + hydrogen	(*i*) Aluminium + sulphuric acid (dil) = Aluminium sulphate + hydrogen (*ii*) Aluminium + sulphuric acid (conc) = Aluminium sulphate + sulphur dioxide + water	No action
Iron	It dissolves in dilute and strong acid Iron + hydrochloric acid + = Iron chloride + hydrogen	(*i*) Iron + sulphuric acid (dil) = Iron sulphate + hydrogen (*ii*) Iron + sulphuric acid (conc) = Iron sulphate + sulphur dioxide + water	Iron + nitric acid (dilute and strong), = Iron nitrate + oxide of nitrogen + water

Metal	*Hydrochloric acid*	*Sulphuric acid*	*Nitric acid*
Lead	No action	Dilute acid has no action, Lead + sulphuric acid (hot and strong) = Lead sulphate + sulphur dioxide + water	Lead + nitric acid = Lead nitrate + oxide of nitrogen + water
Tin	Slow action by dilute acid Tin + hydrochloric acid (hot and con.) = Tin chloride + hydrogen	No action with dilute acid Tin + sulphuric acid (hot and strong) = Tin sulphate + sulphur dioxide + water	Pure nitric acid has no action on tin Dilute nitric acid + tin = Tin nitrate + ammonium nitrate + water Nitric acid (strong) + tin = Metastannic acid + oxide of nitrogen + water
Magnesium	Magnesium + hydrochloric acid (dil.) = Magnesium chloride + hydrogen	Magnesium + sulphuric acid (dil) = Magnesium sulphate + hydrogen	Magnesium + nitric acid (dil.) = Magnesium nitrate + hydrogen.

Action of air on metals. Most of the metals if left exposed to ordinary air are covered with a thin layer of oxide or oxide and basic carbonate due to the action of oxygen and carbon dioxide in air. This action of air on metals is called rusting though commonly used in the case of iron only. In the case of magnesium, aluminium, zinc, tin, lead and copper, the action is very slow in the ordinary air but it becomes rapid when the metals are heated. They form their corresponding oxides. Sodium and potassium are at once tarnished in the ordinary air Gold and silver are not affected

by ordinary moist air or even on heating. Air containing sulphur compounds stains silver. Copper if left exposed for a long time is covered with a green film of carbonate—verdigris. Lead when heated is converted into litharge and red lead. Mercury when heated is converted into red oxide of mercury. In some cases, *e.g.*, in aluminium the layer of oxide formed protects the metal below from further oxidation but in some metals, *e.g.*, iron, the whole of the metal crumbles, and corrodes. In order to prevent rusting of iron the surface is covered by a paint or enamel. It is in some cases tinned or galvanized. It has been seen that iron rusts only in moist air and that dry air or water alone cannot oxidize iron.

Action of water on metals. Metals may be divided into three groups according as they act on water at the ordinary temperature or act when heated or do not act at all. Sodium potassium and calcium decompose water even at the ordinary temperature, combine with oxygen and liberate hydrogen. In the case of potassium the hydrogen burns. Magnesium Zinc, iron, aluminium decompose water at the boiling temperature or in the form of steam. Copper, mercury, silver, and gold do not act upon water at all.

QUESTIONS

1. What do you undertand by the term rusting of metals? Give some common examples. How will you protect iron from rusting?

2. Prove experimentally that both water and air are necessary for the rusting of iron.

3. What is verdigris? How is it formed? What is the use of the thin layer of oxide on the surface of aluminium?

4. Explain the chemical action which takes place when

(1) Hydrochloric acid acts on zinc, (2) nitric acid acts on copper, (3) sulphuric acid acts on common salt, (4) sodium acts on water.
(*P. U.*, *M. & S. L. C.* 1916).

5 Describe experiments to show the action of —

(1) hydrochloric acid on copper, (2) sulphuric acid on lead (3) nitric acid on mercury, (4) plants on carbon dioxide

State clearly the condition under which the above reactions take place and the products of reactions formed in every case

(P U, M & S L C, 1927)

6 Write what you know about the action of —

(a) sulphuric acid on lead, (b) nitric acid on silver (c) hydrochloric acid on copper, (d) water on zinc ?

State clearly the condition under which the above reactions should take place and the products of reaction obtained in each case

(P U, M & S L C, 1932)

7 Describe the action of dilute hydrochloric acid on the following metals—iron, zinc, sodium, aluminium and lead

8 Give the action of water on sodium, potassium, magnesium and aluminium Name some metals which do not decompose water at all

9 Describe the action of dilute and strong sulphuric acid on zinc, mercury, copper, aluminium and iron

10 Describe the action of dilute nitric acid on tin and magnesium and strong nitric acid on aluminium, iron, copper and lead

11 What happens when the following metals are heated strongly in air, copper, zinc, tin, lead and mercury ?

12 What happens when

(a) steam is passed over red hot magnesium,

(b) potassium is thrown in water,

(c) zinc acts upon caustic soda ?

TEST PAPERS

(Taken from University Papers.)

PAPER I

1 Describe a method of distilling water. What impurities would be removed from the muddy water of a pond (*i*) by distillation, (*ii*) by filtration ?

2 How would you prepare crystals of a soluble substance (nitre) given to you in the form of a solid ?

Explain how crystallization may sometimes be used as a means of purifying a substance.

3 What properties must a substance possess to be described as an acid ? Give any examples you can of acids being formed from oxides.

4 What is the composition of (*i*) quicklime (*ii*) slaked lime ? What chemical changes take place when slaked lime is boiled with a solution of sodium carbonate ?

5 Describe exactly the result of heating (*i*) red oxide of mercury, (*ii*) Potassium nitrate (*iii*) lead nitrate (*iv*) ammonium nitrate, (*v*) sodium bicarbonate, (*vi*) blue vitriol.

6 Under what conditions does water act upon the following substances, (*i*) iron (*ii*) sodium (*iii*) charcoal, (*iv*) sulphur dioxide ? What products are formed in each case ?

7 Give an account of the occurrence of aluminium in nature. Describe its chief physical and chemical properties. Give the uses of the metal. What are its useful alloys ?

PAPER II

1 What do you understand by an element, a mixture or a compound ? How would you classify the following ? Diamond, ice-cream, chalk, sulphur, gunpowder, glass, air, and brass ? Give reasons for your answer.

2 Describe the preparation and properties of hydrogen. Sketch the apparatus used.

3. Draw up a table showing the action of (a) air, (b) water, (c) dilute sulphuric acid, (d) strong hydrochloric acid (e) dilute nitric acid on the following metals sodium, aluminium, iron, tin, lead

4 What is a voltameter ? Explain how it is used to determine the composition of water by volume

5 Given the following liquids (i) aqueous solution of sodium chloride, (ii) dilute hydrochloric acid, (iii) dilute sulphuric acid (iv) water—how would you identify each ?

6 Describe experiments to prove that —

(i) A compound is formed when iron and sulphur are heated together in a tube

(ii) A piece of marble and a piece of chalk are chemically the same

(iii) Air contains water vapour

(iv) Substances become heavier on heating in air

(v) Lime juice contains an acid

(vi) Air contains a gas that will not support burning

7 Describe the composition, properties and uses of the following Washing soda, green vitriol, sodium amalgam, brass, gypsum, tin plate

PAPER III

1 Describe two experiments by which you can show that about one-fifth of the air is a gas which is chemically more active than the remainder

2, Explain the term 'water of crystallization' How may its presence in crystals be shown ?

Name two substances whose crystals contain water of crystallization What is the cause of the appearance of a white powder on the surface of some crystals after exposure to the air ? Name two deliquescent substances and explain the meaning of the term

3 Describe how nitrogen is obtained from the air and carry out experiments to show its properties.

4 Define the terms 'acid,' 'base,' and 'salt.' What are the different classes of salts? In what different ways can salts, in general, be obtained? Give illustrative examples.

5 State the conditions which may influence the rate of a chemical action, and give examples to illustrate the effect of each.

6 Explain with illustrations the chief properties by which metals may be distinguished from non-metals.

7 Describe the preparation and properties of nitric acid.

PAPER IV

1 How is a mixture distinguished from a compound? Give an example to illustrate your answer.

2 Describe briefly three methods of preparing oxygen. Give its chief properties.

How would you remove from oxygen traces of (*a*) carbon dioxide (*b*) water-vapour?

3 Describe experiments to illustrate :—

(*i*) The loss of water of crystallization by a salt on heating.

(*ii*) A metal burning in a gas.

(*iii*) Decrepitation.

(*iv*) Bleaching.

(*v*) Sublimation.

(*vi*) Precipitation.

Describe the structure of a candle flame. When a wax candle burns in air what chemical changes take place? How would you show the principal products of combustion experimentally?

5 What is meant by hard and soft water? Distinguish between temporary and permanent hardness of water. What is the hardness of water due to?

Describe the various methods by which hardness of water may be removed.

6. Explain the term 'allotropy.' What are the allotropic forms of sulphur? Describe the preparation and properties of each form.

7. What are the distinguishing characters of cast iron, wrought iron and steel? Give the uses of each.

PAPER V

1. What reasons are there to prove that air is a mechanical mixture and not a chemical compound?

2. Distinguish between (*a*) chemical and physical changes, (*b*) elements and compounds, (*c*) saturated and unsaturated solutions, (*d*) oxidation and reduction, (*e*) alloy and amalgam. Give examples.

3. Give the chief properties and uses of the following metals: Copper, mercury, lead and tin.

4. Describe fully the manner in which the breathing of animals and plants effect the composition of the air.

Carbon dioxide is passed in lime water (1) in small quantity, (2) in excess. Describe the changes that take place. What happens when the solution in (2) is boiled?

5. Give a full account of the allotropic forms of carbon.

6. Describe the preparation, properties and uses of sulphur dioxide? Give experiments to show the reducing properties of the gas.

7. Starting from the metallic sodium, how may the following be prepared: (*a*) caustic soda, (*b*) sodium carbonate, (*c*) sodium bicarbonate, (*d*) sodium sulphate?

What is the action of heat on (*c*) and of water on (*a*)?

PAPER VI.

1. Four jars are given to you containing oxygen, hydrogen, nitrogen and carbon dioxide. By what experiments would you distinguish these gases from one another?

2. Explain the following terms giving examples.

(i) Chemical decomposition, (ii) chemical combination, (iii) catalytic agent, (iv) oxidizing agent, (v) chemical affinity.

3 How is phosphorus obtained from bones? What are its two varieties? Compare the physical and chemical properties of these two forms. How will you show experimentally that these two forms are chemically identical?

4 Describe a method and sketch the apparatus for preparing carbon dioxide gas. How can this gas be converted into carbon monoxide?

Compare the physical and chemical properties of these two gases.

5 What is meant by the solubility of a salt?

Describe with experimental details, how you would determine the solubility of a solid in water.

20 c.c. of saturated soution of epsom salt in water at 20° C weighed 27.4 gms. After evaporating the solution 7.2 gms. of the salt remained. Calculate the solubility of epsom salt in water at this temperature.

6 What different substances can you prepare directly from a supply of sulphuric acid, manganese dioxide, sodium, chlorine, and potassium nitrate.

Explain the terms (1) acid, (2) base, (3) salt, (4) solute, and (5) crystal. Give examples of each.

7. Describe the composition, properties and uses of the following (a) washing soda (b) graphite, (c) lime, (d) blue vitriol (e) brass.

APPENDIX A.

LAWS OF CHEMICAL COMBINATION.

All chemical reactions are governed by certain laws called Laws of chemical combination. The most important of these are briefly given below :—

1. Law of conservation of Mass which states that in a chemical change the total mass of the substances undergoing the change remains the same. This is also known as the Law of Indestructibility of Matter.

For an experimental illustration of this Law the student is referred to Expt. 10 and 11 in Chapter I.

2. Law of constant composition which may be stated as follows :—

A chemical compound always contains the same elements combined together in the same proportion by weight.

The truth of this law can be tested in several ways. Take the example of pure water We have already seen that pure water is made up of hydrogen and oxygen combined together in the ratio 1 : 8 by weight. Pure water whether prepared as in Exp. 65 or obtained from any source in nature is always found to contain hydrogen and oxygen in the ratio 1 : 8 by weight. Similarly copper sulphate from whatever method it is prepared always contains, copper 63 parts, sulphur 32 parts and oxygen 64 parts, by weight and so on.

3 The Law of Multiple proportions which may be stated thus :—

If two elements combine to form more than one compound, the weights of one of the elements which combine with a fixed weight of the other, are in a simple ratio to one another. Thus carbon combines with oxygen in two proportions to form two different compounds, carbon monoxide, and carbon dioxide It can easily be shown that for a fixed weight of oxygen in them there is twice as much carbon in the monoxide as in the dioxide. Here the simple ratio is 2 1.

Structure of matter By suitable means all kinds of matter can be divided into small particles Stones can be reduced to dust so fine that it can float on water A lump of iron can be divided into iron filings Liquids can be divided into tiny drops. These observations lead us to the conclusion that matter is made up of very small particles, called molecules. The molecules can be further divided into smaller particles, known as atoms. Scientists have shown that there are 92 different kinds of atoms. Different combinations of atoms give rise to molecules in the form of which all matter in the universe exists.

SYMBOLS, FORMULAE AND EQUATIONS

Symbols To express the nature of a chemical change, chemists use certain symbols. For each element there is a symbol It consists of the first letter or the first letter together with another letter of the English name of the element or sometimes of its Latin or Greek name. As for example, Hydrogen is denoted by H., Zinc by Zn., Potassium by K (from its Latin

name kalium). Each symbol denotes two things (i) name of the element, and (ii) the atomic weight of the element. For instance S denotes the element sulphur and also one atomic weight of sulphur, *i.e.*, 32 parts of sulphur. A list of the common elements considered in this book is given below :—

(Atomic weights given are only approximate)

Element.	Symbol.	Atomic weight.
Aluminium	Al.	27
Antimony (stibium)	Sb.	122
Arsenic	As.	75
Barium	Ba.	137
Bismuth	Bi	209
Bromine	Br.	80
Calcium	Ca.	40
Carbon	C	12
Chlorine	Cl.	35
Cobalt	Co.	59
Copper (cuprum)	Cu.	63
Florine	F.	19
Gold (Aurum)	Au.	197
Hydrogen	H.	1
Iodine	I.	127
Iron (Ferrum)	Fe.	56
Lead (Plumbum)	Pb.	207
Magnesium	Mg.	24
Manganese	Mn.	55
Mercury (Hydrargyrum)	Hg	200
Nickel	Ni.	59
Nitrogen	N.	14
Oxygen	O.	18
Phosphorus	P.	31

Platinum	... Pt	[illegible]
Potassium (Kalium)	... K	39
Radium	.. Ra.	225
Silicon	.. Si.	28
Silver (Argentum)	... Ag.	108
Sodium (Natrium)	.. Na.	23
Sulphur	... S.	32
Tin (Stannum)	. Su.	119
Zinc	... Zn.	65

Formula By writing chemical symbols alongside each other we can easily express what elements any given compound contains In fact a formula is a group of symbols which stands for one molecule of a substance and expresses its atomic composition. The number of atoms of an element present in a molecule of a substance is expressed by a small figure written just after the symbol. Thus H_2O stands for a molecule of water which is made up of two atoms of hydrogen, and one atom of oxygen. Similarly one molecule of nitric acid HNO_3 contains one atom of hydrogen. The chemical formulae of the chief compounds used in this book are given in the following table :—

Acids.	Hydrochloric acid	HCl
	Sulphuric acid	H_2SO_4
	Nitric acid	HNO_3
Alkalies.	Caustic soda	$NaOH$
	Ammonium Hydroxide	NH_4OH
	Caustic Potash	KOH
	Calcium Hydroxide	$Ca(OH)_2$

Gases.	Ammonia	NH_3
	Sulphur dioxide	SO_2
	Sulphur trioxide	SO_3
	Carbon dioxide	CO_2
	Carbon Monoxide	CO
Oxides.	Manganese dioxide	$Mn\ O_2$
	Phosphorous pentoxide	P_2O_5
	Calcium Oxide	$Ca\ O_2$
	Iron Oxide	$Fe\ O_2$
	Copper Oxide	$Cu\ O$
	Zinc Oxide	$Zn\ O$
	Mercury Oxide	$Hg\ O$
Chlorides.	Sodium Chloride	$Na\ Cl$
	Ammonium Chloride	$NH_4\ Cl$
	Potassium Chloride	$K\ Cl$
	Silver Chloride	$Ag\ Cl$
	Barium Chloride	$Ba\ Cl$
	Mercurous Chloride	$Hg\ Cl$
	Mercuric Chloride	$Hg\ Cl_2$
Nitrates.	Silver Nitrate	$Ag\ NO_3$
	Cupric Nitrate	$Cu\ (NO_3)_2$
	Sodium Nitrate	$Na\ NO_3$
	Barium Nitrate	$Ba\ (NO_3)_2$
	Potassium Nitrate	$K\ NO_3$
Sulphates.	Sodium Sulphate	Na_2SO_4
	Barium Sulphate	$Ba\ SO_4$
	Magnesium Sulphate	$Mg\ SO_4$
	Copper Sulphate	$Cu\ SO_4$
	Calcium Sulphate	$Ca\ SO_4$
	Zinc Sulphate	$Zn\ SO_4$

Carbonates	Calcium Carbonate, (Chalk)	$Ca\,CO_3$
	Magnesium Carbonate	$Mg\cdot CO_3$
	Sodium Carbonate	$Na_2\,CO_3$
	Sodium Bicarbonate	$Na_2\,HCO_3$
	Ammonium Carbonate	$(NH_3)_2\,CO_3$

Chemical Equations. Chemical changes are represented by what are known as chemical equations. The substances taking part in a chemical reaction are placed on the left hand side and the products of the reaction are placed on the right hand side, the various substances on each side being connected by the sign +. Thus water is formed by the combination of oxygen and hydrogen, and this is represented by the equation.

$$2H_2+O_2=2H_2O.$$

The above equation denotes that two molecules of hydrogen combine with one molecule of oxygen to form two molecules of water.

The following chemical reaction further illustrates the use of chemical equations.

Potassium chlorate when heated produces potassium chloride and oxygen.

$$2K\ Clo_3=2K\ Cl+3O_2.$$

The prepration of hydrogen is represented by the equation,

$$Zn+H_2SO_4=Zn\ SO_4+H_2.$$

Again the action of sodium on water is represented by $2Na+2H_2O=2Na\ OH+H_2$.

The neutralization of sodium hydroxide (caustic soda) by hydrochloric acid is expressed as

$$Na\ OH + H\ Cl. = Na\ Cl + H_2O$$

A chemical equation must be correct both chemically and arithmetically.

In other words a chemical equation must express the nature and the composition of the substances taking part in a chemical reaction and also denote their amounts Thus the equation $Zn + 2\ H\ Cl. = Zn\ Cl_2 + H_2$ tells us two things.

(*i*) The nature of each substance taking part in the ractions ; and

(*ii*) 65 parts by weight of zinc react with 72 parts by weight of hydrochloric acid to produce 135 parts by weight of zinc chloride and two parts by weight of hydrogen.

There are a few points to be remembered about the use of figures in connection with chemical symbols. The formula Co_2 denotes 44 parts of carbon dioxide. Suppose we wish to write a formula denoting 132 parts of carbon dioxide. This may be done in two ways (*i*) C_3O_6 or (*ii*) $3\ CO_2$, the number prefixed to the formula applying to all the symbols in it. The second method is always used because it keeps the formula as simple as possible. Again the formula $H_2\ O$ stands for 18 parts of water ; $2\ H_2O$ stands for 36 parts, $3H_2O$ for 54 parts, and so on It must be remembered that fractional numbers are never used in chemical formulae and equations.

APPENDIX B

SCIENTIFIC HOBBIES.

Of late a good deal of stress has been laid on the subject of vocational training in schools and colleges. Several subjects have been suggested with their syllabus by various authors. At present it is not possible for any one school to take up more than a few hobbies. Scientific hobbies on a small scale can be successfully carried on in a school laboratory, as they require simple apparatus, little capital and recurring expenses. These hobbies beside effecting some saving in our daily expenses, are sure to inculcate in us an interest and enthusiasm for small scale industries. A brief outline of some of the simple, cheap, useful and interesting hobbies is given below.

1.—INK MAKING.

1. *Blue Black Fountain pen ink* :—

Galls (Maju) ..	.. ½ seer.
Iron Sulphate (Hira kasis)	.. ⅛ seer.
Gum Arabic (Gond Kekar)	.. ⅛ seer.
Soluble Blue T. ..	.. ¼ lb.
Water ..	.. 12½ seers.
Carbolic acid ..	.. ½ oz.

Process.—Soak the crushed galls in 2 seers of water and let them ferment for a fortnight. Decant

[illegible] the liquid and repeat the process twice with the residue. Collect the Gall extract in a big jar. Dissolve powdered Gum Arabic and Iron sulphate separately in a seer of water, filter and then add the solutions to the gall extract.

Now add soluble Blue T, stir well and make up the solution to 12½ seers. Leave it for a day or two filter and then add to it carbolic acid.

Blue Black Ink Powder or Tablets.—

Logwood extract	..	. 100 parts
Potassium Chromate (*Lal Kahi*)		.. 1 part.
Gum	..	... 20 parts.
Indigo Carmine (*Nil*)	..	... 20 parts

Powder and Mix.—For tablets press under a tablet making machine.

2. *Red Ink.*—

Gall extract (weak)	.	.. 4 seers
R. R. R. Red Colour		.. 2 chhatanks.
Gum Arabic	.	... 2 chhatanks.

Process.—Dissolve the Dye and Gum Arabic in weak solution of gall extract filter and bottle.

3. *Green Fountain Pen Ink.*—

Eosin	...	... 2 oz.
Gum	...	... 2 oz.
Spirit	...	... 20 oz.
Distilled water	...	.. 100 oz.

Dissolve the colour in the spirit with the aid of gentle heat Then in a separate vessel dissolve the gum in water and filter. Finally heat the gum solution and as soon as it begins to boil pour in the colour solution in a thin stream with continuous stirring.

Reference Book.—Manufacture of Ink—By Industry Book Depot, Calcutta.

2 Soap making

(a) *Washing Soap*—This can be prepared by the following cold process—

Mohua Oil (Mowa ka tel)	1 part
Cocoanut oil (Naryal ka tel)	1 part.
Tilli ka Tel	1 part.
Wheat Starch—Maida	1 part.
Soda lye (35°)	2 parts.
Water	$1\frac{1}{4}$ part.
Salt	$\frac{1}{4}$ part.

Procedure.—Melt and mix well the oils and wheat starch. Prepare the caustic soda lye by dissolving caustic soda in water and allow it to cool. Add salt to water and mix it with soda lye. Add soda lye salt mixture to the oil mixture. Stir the mass well with a wooden rod, till it forms a uniform pastry mass. Do not stir too long or the mixture will separate again. Now pour off the pasty soap in wooden moulds and allow it to cool overnight, then remove and cut it into cakes of desired sizes.

N.B.—One seer caustic soda of 77° in $2\frac{1}{2}$ seers of water makes soda lye of 35°.

[illegible]

Cocoa-nut oil	.. ..	16 seers.
Mohua oil		12 seers
Castor oil	..	$2\frac{1}{2}$ seers
Caustic soda lye $37\frac{1}{2}^{\circ}$		15 seers
Water	.. .	5 seers

Heat the oils in an iron vessel of sufficient capacity to a point when it produces smart cracking sounds on putting a few drops of water into it, and a little sprinkling of water produces foam. Now mix the lye slowly and stir it thoroughly. Put out the fire. If the oils were properly heated the mixture will boil in a few minutes. When it calms down add to it 5 seers of water and stir it thoroughly. It will remove all defects from the soap. To impart a pale yellowish tint like that of sunlight soap add a bit of brown colour into the oil before the addition of the lye. Now frame, cool and then cut into bars

(*b*) *Toilet Soap.*—

Mohua oil	. ..	$2\frac{3}{4}$ seers.
Cocoa-nut oil	. .	$2\frac{3}{4}$ seers.
Caustic soda lye 38°	. .	3 seers
(One seer of caustic soda in $2\frac{1}{2}$ seers of water).		
Water	. ..	3 chhataks

Heat the oils to a point when the mixture produces smart crackling sound on putting a few drops of water into it. Now mix the lye slowly and stir it thoroughly. Then put out the fire. If the oil is properly heated it will boil in a few minutes. When it calms down add to it 6 chhataks of water and stir it thoroughly. Now

add the desired colour, perfume and then pour it into frames to cool.

Precautions.—Caustic soda should not be touched with hand as it is corrosive. Its solution should be prepared in iron vessels.

For full information regarding various kinds of soaps the reader is advised to read the following books.—

1 Soap making by the cold process—published by the Government Printing Works, Madras.

2 Manufacture of Soap—published by the Industry Book Depot, Calcutta.

3 Soap making (saban sazi) in Vernacular—Published by Bharat Soap Company, Lahore.

3:—Washing & Stain Removing:

1. *Washing*—Our fine fabric clothes are often spoiled by washer-men due to their ignorance or carelesseness The charges of dry cleaning are so high that everybody cannot afford to get his clothes dry-cleaned, however, these clothes can be cleaned at home without much labour and cost.

Silk garments—Take some Lux or sunlight soap and dissolve it in about a seer of hot water in an enamelled 'chilimchi'. Add sufficient water to soak the cloth in it and rub it well. Repeat this process again and then wash the cloth in cold water, squeeze and dry.

Silken clothes can also be washed like above in [illegible] a teaspoonful of citric acid and gum solution. To give lustre the clothes are ironed in semi dry condition.

Woollen clothes.—1. Shake some powdered soap nuts (retha) in sufficient hot water. Wet the woollen cloth with ordinary water and then wash it well in soap nut solution. If some spots are left then repeat the process with fresh solution. Now wash it with cold water, squeeze, dry and iron

2. Prepare soap jelly, by dissolving 1 part of sunlight or any neutral soap in 3 parts of water. Take a cupfull of this jelly and add to it $\frac{1}{4}$ oz. of liquor ammonia. Dissolve this in sufficient warm water and then wash in it the wet woollen cloth. If some spots are left then add more of soap jelly and work it well. Now wash it with cold water and in a weak solution of acetic acid or citric acid. Wash again with ordinary water. Squeeze. dry in shade and iron.

The following recipes of Dry cleaning fluids will be found useful in washing silk and woollen clothes.

Carbon tetrachloride ..	. 3 gals.
Deodorized cleaner's naphtha	.. $1\frac{3}{4}$,,
Benzol ..	.. 24 oz.
Chloroform ..	.. 2 ,,

Or

Diglycol oleate ..	.. 1 oz.
Water ..	.. 1 oz.

This solution is then dissolved in a solution of

Butyl cellosolve ..	.. 1 oz.
Isopropyl alcohol ..	.. 10 oz.
Carbon tetrachloride ..	.. 14 oz.

[redacted]

Soap	3	parts.
Ammonia	2	,,
Glycerine	1	,,
Methylated spirit	2	,,
Water	50	,,

Dissolve weighed quantity of soap in boiling water, cool and then add the ingredients. Keep the mixture in stoppered bottles and use for washing after removing the stains from the garments.

Stain Removing.—Stains can be readily removed when fresh. On drying it becomes sometimes impossible to remove them without injury to the fabrics.

While removing a stain make the stained spot of the cloth in a hoop and place immediately under it a pad of blotting paper or muslin. Mark the boundary line of the stain with the reagent and then rub the stain with a little quantity of the reagent. Repeat the process on both sides of the garment till the stain disappears, and then dry the cloth. Avoid the use of water unless absolutely necessary

REAGENTS FOR COMMON STAINS.

1. *Oil*.—From cotton—Wash with 10% solution of soda ash or soap solution.

From silk.—Wash with cold solution of ammonia.

From wool. Wash with hot solution of ammonia or Benzene.

[illegible] the spot well with carbon [illegible] turpentine oil. Use the hot liquid if the stain persists. Then wash with soap and ammonia solution. To remove vaseline or machinery grease treat the stain with petrol and proceed as above

3. *Ink (Blue Black)*.—Cotton.—Treat the spot with 4% oxalic acid solution, squeeze and then treat with 5% of Sodium hydrosulphite. If stain persists use hot solutions

Silk—First boil with 4% oxalic acid solution and then treat with 5% sodium sulpho-oxalate solution Repeat the process till stain is removed.

4 *Turmeric*.—Cotton.—Wash with hot solution of soda ash or soap.

Silk.—Wash with dilute ammonia and then with water. Now treat it with 1% potassium permanganate solution squeeze and then put it in 8% sodium sulphite solution. Wash with water and allow to dry.

5. *Blood*. Wash the stain with 5% oxalic acid solution, then with 1% potassium permanganate solution and finally with 8% sodium bisulphate solution. Blood stains on silk garments are difficult to remove.

6. *Betel (Pan) Stains*.—Wash with mustard (kanjee) water.

7. *Paint or varnish*.—Wash with turpentine oil or spirit, or week solution of sodium bisulphite.

8. *Boot Polish St*[illegible] of Turpentine oil.

9. *Iron Rust*—Rub the stain with 5% hot solution of oxalic acid, then treat it with 1% solution of bleaching powder and acetic acid. Wash with water and dry. In case of silk or wool garments wash the stain in oxalic acid and 5% sodium bisulphite solutions.

10. *Fruit Juices.*—Wash with water and then proceed as in 9 above.

11. *Iodine.*—Remove with spirit or 5% Sodium thiosulphate (Hypo) solution.

12. *Silver nitrate*—Wash with 5% Potassium cyanide solution, then with soap water.

13. *Tea, Coffee or Milk.*—Work the stain in 5% borax solution

4 Scented & Medicated Hair Oils.

Note—(*Names are in vernacular.*)

'Mendhi' Leaves	2	chhataniks.
'Kali Chaie'	1	"
Tej Pat	1	"
Hau Ber	1	"
Panri	½	"
Sandal Powder	2	"
Brahmi Buti	1	Tola.
Triphla (Harar, Behera & Anwala	6	chhataniks.
Oil (Tilli oil or Sesame oil)	2	seers.
Water	1	seer.

[illegible] from the above ingredi- [illegible] over night. Now add the [illegible] devoid of water. Filter oil and add [illegible] green for anwala hair oil or ratan [illegible] oil. Then add to it chyper, anwala, rose or moulsari scents as required.

5. Cosmetics.

Vanishing Cream.—

Stearic acid	...	...	8 oz.
Water ...	...	...	56 oz.
Borax (Sohaga)	...	...	1 oz.
Potassium hydroxide (Caustic potash)			$\frac{1}{2}$ oz.
Glycerine ...	...	...	2 oz.

Melt stearic acid in 40 oz. of warm water. Dissolve borax and coustic potash in 16 oz. of water, warm and then mix in it stearic acid water. Stir the mixture well and then add to it glycerine and the required scent.

Cold Cream.—

White Wax ...	...	...	$\frac{1}{2}$ oz.
Spermaceti ...	...	...	1 oz.
Oil of almonds (Badamon ka tel)		...	8 oz.
Rose water ...	...	...	4 oz.

Melt the wax, spermaceti and oil of almonds in an earthen vessel and when nearly cold stir in gradually 4oz. of rose water.

Brilliantine.—

White vaseline	[illegible] lb.
Olive oil (Tel zetoon)	2 dr.
Yellow vasline	1 dr.
Hard paraffine	1 dr.
Scent rose	6 masha

Melt vaseline, paraffin and oil, add scent and bottle.

Pomade.—

Spermaceti	2 oz.
Castor oil	4 oz.
Alcohol	4 oz.
Oil of bergamotte	1 dr.
Oil of neroli	30 mins.
Oil of cloves	10 mins.

Melt the spermaceti, add the castor oil and then the alcohol in which the essential oils have been previously dissolved. Fill in pots and allow to cool without stirring.

Face Powder.—

Zinc white (jist)	100 grams.
Talc (soap stone)	200 ,,
Violet root powder	200 ,,
Rice Starch	500 ,,
Extract of jasmine	15 ,,
Bergamotte oil	0·5 ,,
Rose oil	0·6 ,,
Ylang-ylang	1·0 ,,
Musk tincture	5 ,,

Q. 5. (*a*) How may metals be distinguished from non-metals ?

(*b*) What is an alloy and in what respects does it differ from the metals of which it is formed ? Name three alloys and the metals of which each is composed.

Q. 6. (*a*) Describe an ordinary voltaic cell. (*b*) Explain clearly why such a cell gives a weaker current than a Bunsen's cell ; (*c*) Is it absolutely necessary to use a zinc plate in all kinds of cells ?

Write an essay on the uses of an electric current

Q. 7 (*a*) State the important properties of a magnet.

(*b*) Can a copper wire behave like a magnet ? If so, under what condition ?

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Q. 1 What do you understand by the terms "Chemical combination", "Chemical decomposition" and "Chemical affinity ?" Give examples to illustrate your answer

Or

(*a*) Distinguish between organic and inorganic compounds.

(*b*) What is allotropy ? Illustrate your answer with reference to the allotropic modifications of carbon.

Q. 2. Describe with full practical details, the preparation of hydrochloric acid in the laboratory. What is the action of this acid on :—(*a*) Common salt, (*b*) Manganese dioxide, (*c*) Sodium carbonate and (*d*) Copper ?

Three beakers contain solutions of :—(a) copper sulphate, (b) sodium chloride and (c) potassium nitrate respectively. What experiments would you perform to determine the contents of each beaker?

Q. 3. Describe what happens when roll sulphur is heated in a test tube. What different varieties of sulphur exist, how are they prepared and what are their properties?

Or

How may sulphur dioxide be prepared? Describe its important properties and uses

Q. 4 Describe what happens and name the product or products formed when :—

(*a*) Excess of carbon dioxide is passed through lime water

(*b*) Hydrochloric acid is added to a solution of caustic soda.

(*c*) Iron is dissolved in dilute sulphuric acid.

(*d*) Concentrated sulphuric acid acts on phosphorus pentoxide

Or

Give the composition, properties and uses of the following :—

(*a*) Gypsum : (*b*) Limestone, (*c*) Glauber's salt, (*d*) Nitre.

Q. 5 Enumerate the important properties and uses of any *three* of the following :—mercury ; zinc ; lead ; tin ; and iron.

Or

Describe the preparation, on a large scale, properties and uses of caustic soda.

Q. 6. What is Mariner's Compass? Explain briefly the construction and use of this instrument.

Q. 7. (*a*) Explain what is meant by :—(*i*) electric induction, and (*ii*) electric attraction and repulsion

(*b*) What are conductors and insulators?

To which class do the following belong :— Human body; glass rod: and metal; and wood.

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1. Describe the following processes and give examples of their uses :—

(*a*) crystallization, (*b*) sublimation, and (*c*) distillation.

Or,

Draw a diagram of a candle flame. Describe its different parts. Show the hottest zone.

What is the effect of blowing air into the flame through a blow-pipe?

II. (*a*) Describe, with full practical details, the preparation of nitric acid in the laboratory.

(*b*) What is the chemical action of sulphuric acid on :—(*i*) zinc, (*ii*) copper, (*iii*) sodium chloride and (*iv*) sugar?

Or,

There are four beakers containing solutions of :— (*a*) caustic soda; (*b*) calcium hydroxide; (*c*) hydro-

chloric acid ; and (*d*) nitric acid. What experiments would you perform to determine the contents of each beaker ?

III. Name the usual impurities in the following: (*a*) rain water (*b*) mineral water ; (*c*) hard water ; and (*d*) sea water.

Describe the electrolysis of water.

Or,

How may chlorine be prepared ? Give its important properties and uses.

IV. Describe the chemical action of —(*a*) sodium on water ; (*b*) nitric acid on copper ; (*c*) lime on a solution of sodium carbonate ; (*d*) hydrochloric acid on marble

Or,

Give the composition, uses and properties of the following –

(*a*) potassium chlorate, (*b*) slaked lime ; (*c*) ammonia and (*d*) sulphur dioxide.

V. What do you understand by an element, a mixture, and a chemical compound ? How would you classify the following :—diamond ; ice cream ; chalk ; sulphur ; gunpowder ; and brass ? Give reasons for your answer.

Or,

Write a short essay on the different forms of carbon and their uses.

VI. Describe the various methods of making magnets.

VII. (*a*) Describe the action of electrophorus.
(*b*) How would you set up a Daniell cell ?

University Syllabus on Chemistry—1941-42

Physical and chemical changes. Elements, compounds, and mixtures, metals and non-metals. Chemical combination and decomposition.

Solution, decantation, filtration, evaporation, distillation, Saturated Solution, Crystals, Crystallization, Sublimation, Composition of air, Oxygen. Nitrogen, Impurities of air, Deliquescent bodies, Air a mixture.

Chemical affinity, preparation of Oxygen, its properties and uses. Oxidation, reduction, combustion, Parts of a candle flame.

Hydrogen, its preparation and properties.

Water, its properties and composition by electrolysis. Rain, spring, mineral, and sea-waters ; hard and soft waters. Softening of hard waters

Organic and inorganic compounds. Allotropic forms of Carbon. Coal, breathing. Burning of a candle. Action of plants on CO_2.

CO_2, its preparation, properties and uses. Limestone. Lime and slaked lime. Forms of Calcium Carbonate and Calcium Sulphate.

HCl, its preparation and properties, Chlorides. Sulphur. Its varieties, effects of heat upon it in a closed vessel and air, SO_2, and its properties. Sulphuric Acid, its properties and action on metals. Sulphates, Phosphorus—red and yellow, matches.

Nitric acid, its preparation, properties and uses Nitrates, Distinction between HCl, H_2SO_4 and HNO_3. Ammonia and its Properties.

Ores and Metals. Alloys and Amalgams. The more important salts of Sodium and Potassium.

Properties of the following metals: Copper, Mercury, Silver, Zinc, Lead, Tin, Iron and Aluminium.

Practical Chemistry.—Acquaintance with simple chemical manipulation, as solution, filtration, decantation, crystallization, distillation.

To fit up an apparatus to demonstrate the combination of oxygen of the atmosphere with iron.

Cork boring; cutting, bending and drawing out of glass tubing and glass rod.

The preparation and properties of oxygen and hydrogen

The distinction between acids and alkalies.

To neutralize an acid with an alkali *vice versa.*

A study of the zones in a candle flame.

The products of combustion as illustrated by the burning of a candle.

To distinguish between hard and soft waters and to soften hard water.

The action of heat on coal to show the production of coal gas.

The preparation and properties of carbon dioxide, hydrochloric acid and nitric acid.

Test for hydrochloric acid, sulphuric acid and nitric acid.

The action of heat on sulphur.

To prepare monoclinic and plastic forms of sulphur.

1910

I. Describe the preparation, the properties and the uses of oxygen.

Or

Write what you know of the oxygen cycle in nature, and explain oxidation and reduction ?

II. How is carbon dioxide prepared ? What are its properties and uses ?

Or

What is the difference between :—

(*a*) Limestone. (*b*) Chalk, (*c*) Lime, (*d*) Slaked lime, (*e*) Gypsum, (*f*) Plaster of Paris, (*g*) Precipitated chalk, and (*h*) Mortar ?

III. Describe the preparation, the properties, and the uses of HCl

Or

What is the action of HCl on :—

(*a*) Zinc carbonate, (*b*) potassium permanganate, (*c*) sodium sulphite, (*d*) iron filings ?

IV. Describe at least four chemical properties of :—(*a*) Zinc, (*b*) Aluminium.

Or

What is the difference between :—

(*a*) A physical and a chemical change, (*b*) metals and non-metals, (*c*) amorphous and crystalline, (*d*) red and yellow phosphorus ?

V. Write a short essay on coal, describing its uses and the products obtained from it.

Or

(*a*) How is lead obtained from its ore ?

(*b*) Give the uses of lead

VI. Write a short essay on sulphur and its varieties, and describe the effect of heat on it in air, and in a closed vessel

VII Describe the chemical action of :—

(*a*) Lime on ammonium chloride

(*b*) Potassium on water

(*c*) Manganese dioxide on HCl.

(*d*) Copper filings on concentrated sulphuric acid

(*e*) Sodium carbonate solution on zinc chloride solution.

(*f*) HCl on sodium sulphite.

(*g*) Caustic Soda on zinc

(*h*) Conc. sulphuric acid on sodium chloride